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ABSTRACT

This publication is the third product of the project to significantly update and reorder the Science Syllabus for Middle and Junior High Schools of the State of New York, Department of Education. The understandings, questions, and learning activities in the publication have been enhanced by the inclusion and referencing of program objectives, performance criteria, and process skills which serve to focus student achievement. Environmental, consumer, and career implications are included in the material. Unit titles include: (1) Properties and Changes in Matter; (2) Introduction to Atomic Structure; (3) Chemical Shorthand; (4) Introduction to Organic Chemistry; and (5) The Role of Chemistry in Society. Learning activities deal with, among other things, laboratory safety and skills, physical properties of matter, chemical kinetics and equilibria, and solubility. Some activities require a laboratory facility while others do not. Many activities and concepts could be used in physics and biology instruction as well. (MR)

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Science Syllabus for Middle and Junior High Schools

U.S. DEPARTMENT OF HEALTH, EDUCATION & WELFARE
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The CHEMISTRY of MATTER

block H

SE 022 825

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The University of the State of New York
THE STATE EDUCATION DEPARTMENT
Bureau of General Education Curriculum Development
Albany, New York 12234

SCIENCE SYLLABUS

FOR

MIDDLE AND JUNIOR HIGH SCHOOLS

BLOCK H

THE CHEMISTRY OF MATTER

THE UNIVERSITY OF THE STATE OF NEW YORK

Regents of The University (with years when terms expire)

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FOREWORD

This publication is the third product of the project to significantly update and reorder the Science Syllabus for Middle and Junior High Schools. The project will result in a group of instructional blocks encompassing the content and skills in the biological, physical, and earth sciences appropriate for the grade levels indicated above.

The understandings, questions, and learning activities in the publication have been enhanced by the inclusion and referencing of program objectives, performance criteria, and process skills which serve to focus student achievement. Thus, environmental considerations are infused throughout the content material and the importance of consumer and career implications in a science program is specifically acknowledged.

The objectives and performance criteria are provided in a general outline on pages 1 through 3 which conforms to the content subdivisions in the publication; the process skills are keyed to the learning activities; and the learning activities are keyed to the understandings and the questions.

The development of this syllabus was coordinated within the Bureau of General Education Curriculum Development by Barry W. Jamason, associate. Elizabeth Lamphere, Norwich Public Schools, revised portions of the content material while reviewing the entire manuscript. Edward T. Lalor, Chief, Bureau of Science Education, and John V. Favitta, associate in science education, reviewed all materials and made valuable suggestions for content improvement and accuracy. The cover design was created by Oscar A. Kaufman.

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PROGRAM OBJECTIVES AND PERFORMANCE CRITERIA

The program objectives and performance criteria below are provided as an aid in ordering the direction and synthesizing the outcomes of the instruction related to this block. The objectives and outcomes (performance) described below for this program permit latitude in the selection of content. It is implied that student knowledge, understanding, and skills accrued as a result of content selection and instruction will be demonstrable in the manner, and to the minimum extent described.

Program Objectives	*Performance Criteria*
I. INTRODUCTION	
The student will be able to cite the general safety rules and methods of preventing injury in the laboratory.	Given a situation involving laboratory or home procedure, the student will be able to identify the proper safety considerations involved in carrying out the procedure.
The student will be able to demonstrate proper procedure for the use of laboratory tools such as burners, glassware, and porcelainware.	Given an example of a laboratory or home accident, the student will be able to demonstrate the activity necessary to minimize the possible injury to people and facilities, and identify the errors in procedure or judgment that precipitated the accident.
The student will be able to explain and demonstrate proper procedure for handling laboratory and household chemical substances.	Given a situation involving handling of school laboratory or household chemicals, the student will be able to demonstrate proper handling procedures.
II. PROPERTIES AND CHANGES IN MATTER	
The student will be able to differentiate between the physical properties and chemical properties of matter.	Given a substance, the student will be able to identify its physical and chemical properties.
The student will be able to describe examples of physical changes and chemical changes in matter.	Given a substance, the student will be able to cite an example of a physical change that could be produced in the substance.
	Given an example of the chemical change of a substance, the student will be able to cite the conditions necessary for that change to have occurred.

Program Objectives	*Performance Criteria*
III. INTRODUCTION TO ATOMIC STRUCTURE	
The student will be able to describe a model of the atom and use this model to explain chemical activity.	Given an element, the student will be able to describe an atomic model of that element. Given a compound, the student will be able to describe the molecular model of that compound.
IV. COMMON CHEMICAL CHANGES	
The student will be able to recognize the various types of chemical changes and predict the result of common reactions.	Given a chemical reaction, the student will be able to identify the reaction as direct union, replacement, or decomposition.
V. CHEMICAL SHORTHAND	
The student will be able to demonstrate sufficient acquaintance with symbols, formulas, and coefficients used in chemical equations to interpret them.	Given a chemical reaction in symbolic form, the student will be able to identify the substances represented by the symbols.
VI. COMMON COMPOUNDS AND MIXTURES	
The student will be able to differentiate between compounds and mixtures and the properties that determine their characteristics.	Given a substance, the student will be able to identify it as a compound or mixture. Given a compound, the student will be able to identify it as an acid, base, or salt.
VII. INTRODUCTION TO ORGANIC CHEMISTRY	
The student will be able to cite the differences between organic and inorganic compounds, and name at least two classes of organic compounds.	Given the name or formula of a compound, the student will be able to identify it as organic or inorganic. Given the name or formula of an appropriate organic compound, the student will be able to classify the compound.

Program Objectives	*Performance Criteria*

VIII. THE ROLE OF CHEMISTRY IN SOCIETY

The student will be able to describe the reliance of living things upon chemical processes.	Given a chemical or a chemical process, the student will describe how it is critical to the maintenance of living systems and/or human economic systems.
The student will be able to propose ways of husbanding chemical resources.	Given a chemical or chemical product, the student will describe how it may be conserved, salvaged, properly disposed of, or recycled.

Throughout this program the instructor should be cognizant of the importance of including several related objectives and performance criteria which represent the practical infusion of consumer and career considerations as they relate to science education:

The student will be able to demonstrate the relationship between chemical products and their economic function as available resources.	Given the characteristics and use of a chemical, the student will demonstrate understanding of its significance as one of many resources with a role in our economy.
The student will be able to demonstrate an awareness of the contribution to society which is being made by persons who have elected careers and occupations which are allied with the subject matter.	Given the name of a career or occupation related to the subject matter, the student will describe the contributions to society made by the able execution of such career or occupational activities.
The student will be able to describe the career and occupational opportunities and the knowledge, skills, and attitudes helpful in those careers and occupations related to the subject matter.	Given a career or occupational classification related to the subject material, the student will describe activities carried out in this work.
	Given an occupational role related to the subject material, the student will describe the training and aptitudes required to perform the role.
	Given an opportunity, the student will describe currently practiced work habits, skills, and knowledge that relate to careers implied in the subject material.

PROCESSES

One of the major objectives in all science courses is the development of competence in processes used in science. The learning activities in this publication will assist students in attaining this objective. Each activity is keyed to one or more of the processes listed below. For example, activity (H-1) involves processes [2], [8], and [9]. In general, the processes are listed in order of predominance within the activity and appear to the right of the activity's title.

[1] Observing objects and phenomena

[2] Describing observations

[3] Selecting suitable measuring instruments

[4] Developing skills in using laboratory equipment

[5] Measuring objects and phenomena

[6] Performing laboratory techniques

[7] Estimating measurements within limits of accuracy

[8] Recognizing a problem

[9] Formulating a working hypothesis

[10] Designing appropriate tests of a hypothesis

[11] Designing suitable procedures for experimental tests

[12] Processing experimental data

[13] Organizing data to show a functional relationship

[14] Interpreting observations and data

[15] Interpolating and extrapolating

[16] Evaluating a hypothesis in the light of data obtained

[17] Formulating generalizations based on relationships

[18] Recognizing the need for a theoretical model

[19] Formulating a theoretical model to accommodate knowledge

[20] Designing relationships satisfied by a model

[21] Deducing a hypothesis from a theoretical model

[22] Interpreting and evaluating tests of a model

[23] Producing a revised, refined, or extended model

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
I Introduction	
A Safety in the laboratory	
A laboratory is as safe as the least safe person. It is the responsibility of each pupil to avoid accidents.[1]	*How can pupils promote safety in the laboratory?* (H-1)
1 general rules	
Pupils may expect that teachers will set good examples by observing all safety measures.	*Have pupils assist you in developing a list of safety rules.*
Pupils should read, understand, and follow directions, working quietly, thoughtfully, and efficiently.	
Pupils should maintain strict discipline; talking and moving about the laboratory should be kept to a minimum.	
Pupils should work in the laboratory only when a teacher is present.	
Pupils should be able to answer questions about procedure, demonstrating that they know what they are doing.[2,3]	
2 preventing injuries	
A laboratory is no place in which to eat or drink. Avoid any situation which may lead to any chemical being taken internally since most chemicals are poisonous.[4]	*How can chemicals be prevented from entering the body?* (H-2)
Pupils must protect their clothing and eyes by wearing rubber aprons and goggles or eye shields.	
Wash away with quantities of water any chemical spilled on the skin, clothing, or desk.	*What are the effects of chemicals on clothing?* (H-3)

[1]Safety procedures are a part of all scientific investigations; must be stressed continuously; and must not be isolated into a single teaching unit.

[2]Many accidents can be prevented before they occur. Keep close watch on the accident-prone pupils and the one who has a "home lab."

[3]It is the responsibility of the teacher to see that first aid is given. Have the school nurse or doctor, when immediately present, treat serious injuries. Have the principal or his delegate immediately inform the pupil's parents. Limit treatment to first aid. Further treatment is the parents' responsibility.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Keep the desk, equipment, and hands clean at all times. Mere traces of some chemicals can cause other materials to react or decompose violently.[5]	*How can contamination of residual matter in unclean glassware be demonstrated?* (H-4)
Use only a clean spatula to transfer solid chemicals.	
While pouring liquids, hold the bottle stopper or cork so that liquid will not get on the desk or other materials get on the cork.	*What is the proper procedure for pouring liquids from a reagent bottle?* (H-5)
When observing the odor of a substance, fan the gas toward the nose. Do not put the nose directly over the material. Many substances can irritate the membranes of the respiratory system.[6]	
3 using safety devices	
Know how and when to use the safety devices. Use the fire blanket when a person's hair or clothing is on fire. Use the carbon dioxide fire extinguisher for other fires.	*What safety devices should pupils know how to use?* (H-6)
Report all accidents, even the smallest ones. An injured person may not realize that first aid is necessary.	*How should accidents be reported?* (H-7)
B Use of laboratory tools	
1 burners	
a use of burners	
Bunsen burners, propane or butane burners, alcohol lamps, or a candle may be used as a source of heat in the laboratory. The amount of heat needed determines which source to use.	*Which produces more heat, a candle or a burner?* (H-8)

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[4]The *Merck Chemical Index, an encyclopedia of chemicals and drugs, 8th edition, 1968* published by Merck and Company, Rahway, New Jersey is a valuable reference for toxicity levels, entry routes into the body, and symptoms. It also gives additional information about the chemicals listed.

[5]When preparing oxygen for use in a demonstration, use either the electrolysis of water or the decomposition of a three percent hydrogen peroxide solution with a slight amount of manganese dioxide catalyst. Do not use potassium chlorate to prepare oxygen since the reaction is too dangerous. Traces of carbon, excessive heat, or pressure can cause the salt to explode violently.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
The parts of burners are the spuds or gas openings, the air holes or ports, and the barrels.	
When a proper air-gas mixture is obtained, the gas burns with a pale blue flame and has a cone-like appearance. There is enough air being supplied to the burner for the fuel to be completely burned.[8]	*How does a burner operate and what is the nature of its flame?* (H-9)
The hottest point of the flame is at the tip of the inner blue cone. The hottest part of the flame is called the reducing flame.	*What is a reducing flame?*
The tip of the outer flame is relatively rich in oxygen and is referred to as the oxidizing flame.	*How is a candle flame like a burner flame?* (H-10)
b safety rules for a burner	
Light the match, then turn on the gas. If too much air-gas mixture accumulates, an explosion can occur when a match is lighted.	*What should be done if an air-gas mixture accumulates in an oven or room because the flame goes out or because the gas is turned on accidentally?*[9]
When lighting the burner, keep the face and other parts of the body to one side of the burner. Never reach across a burner; hair and clothing can be set on fire.	
Always apply heat gradually and evenly. Use a small wire gauze under any glassware being heated on an iron ringstand.[10]	
When heating a test tube, point the mouth of the test tube away from yourself and others. As the liquid begins to boil, it may spurt out across the desk and burn someone.	*What is the correct procedure for heating a liquid in a test tube?* (H-11)
Carefully follow all directions regarding the amount of heat to use when heating chemicals. Chemical reactions may get out of control as the temperature increases.	

[6] Do not prepare carbon monoxide, hydrogen sulfide, chlorine, ammonia, hydrogen chloride, sulfur dioxide, or gaseous nitrogen compounds. The general science laboratory is usually not equipped for the safety preparation of quantities of these or other poisonous gases.

[7] Each pupil involved in an accident should be required to fill out the accident form in accord with policy developed by the local school. Not only is the procedure a good safety-teaching device, but a record is available in case it is ever needed. See appendix (H-7) for a suggested form to be used in reporting accidents.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
As a general rule, the rate of a reaction increases rapidly with an increase in temperature.	*Does heat affect the rate of a reaction?* (H-12)
2 glassware	
Pyrex or fire resistant glass is recommended since it expands to a very small degree as it is heated. Therefore, it is not apt to break when direct heat is applied gradually.	*How is a funnel used in filtering?* (H-13)
3 cutting, fire polishing, and bending glass	
Newly cut glass tubes or rods should be fire polished to round the ends and reduce the possibility of injury.[11]	*How is glass tubing cut, fire polished, and bent?* (H-14)
When bending glass, use a wing top on the burner since it is necessary to have a larger amount of glass heated than in fire polishing.[12]	
4 porcelainware	
Evaporating dishes should be placed on a wire gauze or heated over a water bath.	*How are solutions evaporated?* (H-15)
Mortars and pestles should never be heated or used for anything except pulverizing solids. Some solids explode upon being ground.	
Do not mix chemicals by grinding them together in a mortar. They may react violently because of the pressure being applied.	
C Household chemicals	
Most household chemicals are poisonous.	

[8]Relate the gas intake to Bernoulli's principle. As the gas under pressure is released through the small opening of the spud, there is an increase in the velocity of the gas causing a reduced pressure in the gas stream. Air rushes in through the air ports due to pressure differences, and a combustible air-gas mixture passes up the barrel.

[9]Many accidents occur each year when an accumulation of gas in an oven is exploded as someone tries to relight the oven burner. A violent explosion occurs and not only damage to property results but also serious burns or death for the person lighting the match.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Many cleaners, paints, polishes, waxes, and lighter fluids give off toxic vapors which are flammable. Use these substances only in the absence of a flame and in a well-ventilated place.	*What precautions should be taken so household chemicals can be used and stored safely?* (H-16)
Avoid mixing household cleaners together since they may produce poisonous gas.	
Aerosols explode when punctured or heated. Always use them according to the directions.	*Why are aerosols harmful to the environment?*
Know what first aid is necessary if by accident a household chemical gets into your eyes, on your skin, or into your body.	*What is the "universal antidote?"*
Light decomposes some substances. Keep chemicals which come in dark or opaque bottles out of the sunlight. Store them and any other household chemicals in a cupboard well out of the reach of small children and pets.	
Keep all containers of household chemicals tightly closed to avoid spillage and evaporation of toxic solvents.	
II Properties and Changes in Matter	
Matter is anything that occupies space and has mass. Under the influence of gravity, mass has weight.	*How can the different kinds of matter be shown?* (H-17)
Three classes of matter are elements, compounds, and mixtures.	
An element is a substance that contains only one kind of atom.	
Elements are not able to be broken down into more simple substances by ordinary chemical means, but may be broken down by nuclear reactions.	

[10]When heating a test tube containing a liquid, heat the contents evenly by moving the burner flame slowly up and down along the outside of the tube. If heat is applied strongly to the base of the tube, the contents in the bottom of the tube will change from a liquid to a gas and cause the liquid above to spurt from the test tube.

[11]Etching the glass with a file produces strains which reduce the force (stress) required to break or shear the material.

[12]Glass becomes more fluid upon being heated. When fire polishing glass, hold the tubing at an angle, and rotate it so that gravity will not cause the heated end to bend toward the ground.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Most elements are metals.	
A compound is a substance which contains more than one kind of atom in chemical combination.	
Compounds are made up of two or more elements chemically united in a definite proportion by mass so that none of the component elements can be recognized.	
Mixtures contain two or more substances in varying amounts but the individual substances can be recognized because they are not chemically united.	*Can the constituents of elements be identified in their compounds?* (H-18)
A Physical properties	
All substances have special physical properties.	*What are some special physical properties?*
1 phases	
The phase in which matter exists depends upon how close together its particles are and how fast they move.[13]	*Which phase changes absorb energy?* (H-19)
When a change of state occurs, energy is either absorbed or liberated.	
Heat energy can be converted into kinetic energy of motion and/or potential energy of position of the particles.	
Kinetic and/or potential energy can be converted into heat.	*What are some examples of kinetic and potential energy?*
The particles of a solid are relatively close together and move relatively slowly resisting a change of position. A solid has a definite shape and volume.	

[13]Avoid using the term molecule in connection with solids and liquids. Some materials are composed of ions (charged particles) bound together in a crystal lattice or moving about in the liquid state or solution. Nonionic substances can exist as molecules in the liquid or solid state. Refer to an advanced high school chemistry textbook or an inorganic college chemistry text for a discussion of the molecular concepts.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
The particles are farther apart and move more rapidly in the liquid phase than in the solid phase.	
Energy is absorbed when a solid liquifies and is released when a liquid becomes a solid.	
A liquid takes the shape of the container but has a definite volume.	*How is a liquid different than a solid?* (H-20)
The particles are farther apart and move much faster in the gaseous phase than the particles in a liquid or a solid.	
Energy is absorbed when a liquid changes into a gas. Energy is liberated when the gas condenses into a liquid.[14]	
Gases have neither a definite shape nor a definite volume. They assume the shape and volume of the container.	*How is a gas different than a liquid? How is a gas different than a solid?*
2 melting and freezing points	
A pure solid has a definite temperature called the melting point, at which it turns into a liquid.	*Do all solids have the same melting points?* (H-21)
A liquid changes into a solid at a temperature called its freezing point.	
The melting and freezing points are the same temperature for the same substance.	
It is possible for a substance to exist as a solid and a liquid at the same temperature. The melting point of pure water is 0° Celsius at 1 atmosphere of pressure.	

[14]The energy being absorbed or liberated during a change of phase is involved with the particles adjusting themselves to the new conditions necessary for the new phase. Primarily, it is a shift of potential energy since the constant temperature indicates that the average kinetic energy of the particles is remaining constant.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
3 boiling and condensation points	
A pure liquid boils at a definite temperature which depends upon the atmospheric pressure The boiling point of pure water is 100° Celsius at one atmosphere of pressure.	*How can temperature be observed while solids, liquids, and gases undergo phase change?* (H-22)
As the pressure on a liquid increases, the boiling point increases.	*Do all liquids boil at the same temperature?* (H-23)
	How can the information about the boiling and melting points be used? (H-24)
A gas condenses into a liquid at the condensation point.	
The condensation and boiling points are the same temperature under the same pressure.	
It is possible for a substance to exist as a gas and a liquid at the same temperature.	
4 other physical properties	
Certain substances have special characteristics of tenacity, ductility, malleability, elasticity, brittleness, hardness, and color.[15] The properties vary in degree from substance to substance.	*How can some physical properties of metals (tensile strength, malleability, ductility, and hardness) be determined?* (H-25)
B Physical changes	
All matter can be changed physically without any change in composition.	
During a physical change there can be a change in the form or energy of the substance, but there is no change in the composition.	

[15] If a material resists being pulled apart, it has tenacity. Malleable material can be beaten or rolled into thin sheets. Ductile substances can be drawn into fine wires. Brittle material tends to fracture under shock or pressure. Hard materials have the ability to scratch, etch, or cut other materials.

The color of a solid is the same as the color of the light being reflected from it. Some materials appear colored because they absorb certain light waves and reflect others.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
1 separating by magnets	
A magnetic substance can be removed from a mixture by means of a magnet.	*How can a magnet be used to produce a physical change?* (H-26)
2 dissolving substances	
Some substances are soluble; other are not.[16]	*Are substances equally soluble in a specific volume of water?* (H-27)
The degree of solubility of a substance at different temperatures may be shown by a chart or graph.[17]	*How can the solubility of a solid be predicted?* (H-28)
The solubility of a gas increases with an increase in pressure; the solubility of a gas decreases with a decrease in pressure.	*How can a solubility curve be determined? Of what use is it?* (H-29)
	What factors affect the solubility of gases in water? (H-30)
3 filtering, decanting a suspension	
An insoluble substance can be separated from a soluble one by filtering. The insoluble material is left on the filter paper. The soluble material is in the filtrate.	*How can some mixtures be separated?* (H-31)
An isoluble substance may be partially separated from a soluble one by decantation. After the mixture has settled, as much liquid as possible is poured off without disturbing the solid.	*What is the process of decantation?*
4 evaporating solutions	
A soluble substance can be recovered from a solution by evaporation.	*How can a solute be obtained from a solution?* (H-32)

[16] No substance is completely insoluble.

[17] Chemical dictionaries and handbooks will give the solubilities of substances in grams per 100 grams of water. Solubility tables such as the one found in the appendix (H-28) classify substances as soluble, slightly soluble, or insoluble. A graph may also be used to show the solubility of a substance at different temperatures.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
5 changing phase	
A change of phase is a physical change since there is no change in composition.	*How can a change of phase be demonstrated as a physical change?* (H-33)
C Chemical properties	
All substances have special chemical properties. Substances show varying degrees of ability to react with other substances.	*How can differences in chemical activity be demonstrated?* (H-34)
When substances combine with oxygen and noticeable light and heat are emitted, rapid oxidation or burning is occurring.	
Some substances react with oxygen without noticeable light or heat being given off. Such reactions are examples of slow oxidation.[18]	
Some substances react with water; some substances react with acids; and some substances react with bases.	
The gas given off when a metal combines with an acid other than nitric acid is hydrogen.	
D Chemical changes	
Matter can be changed chemically.	*How can chemical changes be recognized?* (H-35)
Chemical change results in the formation of one or more new substances, each having its own properties by which it is recognized. The properties of the original material can no longer be detected since they no longer exist.	

[18]As pupils are exposed to more advanced chemistry, the definition of oxidation will take on the broader meaning: oxidation is a loss of electrons by an atom or ion. Pupils should be informed that the definition is limited and to expect revisions later.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Chemicals react according to a definite combining ratio. If the correct ratio is not used, the excess reagent will not combine and may be identified.	*What is the relationship between the amounts of the chemicals in a reaction?* (H-36)
1 chemical change and energy	
A chemical reaction involves a specific amount of energy before the reaction can occur. This energy is called activation energy. Once started, some chemical changes may produce energy as heat, light, or electricity.[19]	*What is meant by activation energy?*
A chemical reaction which produces energy is called an exothermic reaction.	*How is energy related to chemical changes?* (H-37)
Once started, some chemical reactions need energy in the form of light, heat, or electricity to continue. A reaction which uses energy as it progresses is called endothermic.[20]	*What are some differences between exothermic and endothermic reactions?*
2 changing the rate of reaction	
The rate at which a reaction occurs can be changed.	*How can the rate of a reaction be changed?* (H-38)
A catalyst is a substance which can change the rate of a chemical action without its own composition being changed at the end of the reaction.[21]	*How can the effect of a catalyst be shown?* (H-39)
As the concentration of the materials increases, the rate of the reaction generally increases and may get out of control. For this reason use only small amounts of reacting substances.	*How can we demonstrate the relationship between concentration and rate of reaction?* (H-40)

[19] The light from burning magnesium contains harmful ultraviolet rays. Do not allow pupils to look directly at any vigorous reaction producing light.

[20] The products of exothermic reactions tend to be stable, while some or all of the products formed by endothermic, explosive. Nitrogen compounds are in this latter class.

[21] Some catalysts temporarily enter into the reaction to form complex substances which immediately produce new substances and release or reform the catalyst. Other catalysts do not directly enter the reaction but collect gas molecules to their surfaces, the gas molecules coming into closer contact with each other more rapidly.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
According to the law of mass action, the reaction velocity varies directly with the product of the concentration of the reactants.	*What is the law of mass action?*
As the temperature increases, the rate of the reaction usually increases and may get out of control. For this reason, do not heat chemicals unless told to do so. Then carefully follow all directions for heating.	*How can the effect of increased temperature on the rate of reaction be shown?* (H-41)
The smaller the particles of reactants are, the faster they combine.[22]	*How can it be shown that the surface area affects the rate of chemical reactions?*
Many reactions are safely carried out by using solutions of the reactants.[23]	
III Introduction to Atomic Structure	
Atoms occur singly as the smallest particles of an element that can exist and still retain the physical and chemical properties of the element.	
An atom consists of a nucleus around which one or more electrons may be found.	*How can the existence of atoms and molecules be shown?* (H-42)
An atom can be considered to be mostly empty space.	
For most practical purposes, the entire mass of an atom may be considered to be in the nucleus.	
There is no balance sensitive enough to determine the mass of a single atom. The atomic mass is the mass of an atom as compared with the mass of a carbon atom taken as 12 units.[24]	*How can relative masses be determined?* (H-43)

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[22] Some substances such as copper sulfate and sodium chloride can be pulverized safely. Under no circumstances pulverize any chlorate, ammonium, nitrate, or nitrite compound.

[23] There are other factors such as pressure of combining gases and the nature of the reactants that affect the rate of reaction.

[24] The atomic mass of an atom is the relative mass in comparison to the mass of an arbitrary standard. Atomic mass is frequently, although incorrectly, referred to as atomic weight.

One standard is the 12 units taken as the mass of the carbon-12 isotope.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
A The atom	
1 nucleus	
The protons and neutrons are concentrated in the central part of an atom called the nucleus.	
Protons and neutrons are also called nucleons. The atomic mass number equals the number of nucleons. The proton carries a single positive charge and has a mass equal to one unit.[25]	*Where are the atomic masses and numbers found on the Periodic Table?*
The atomic number of an atom indicates the number of protons in the atom. No two elements may have the same number of protons in their nuclei.	
A neutron has no electric charge; it is neutral. The mass of a neutron is equal to one unit.	
The number of neutrons can be found by subtracting the atomic number from the mass number of the atom.	
Isotopes are atoms of the same element having different numbers of neutrons.	*How are isotopes alike and how are they different?* (H-44)
Isotopes have the same atomic number but different atomic masses due to the different numbers of neutrons.[26]	
Isotopes of the same element have the same atomic number and the same chemical properties.	
2 electrons	
The electron has a single negative charge.	

[25]The mass of a proton or of a neutron is considered to be one unit. An "isolated" proton actually has a mass of 1.0081 units, and an "isolated" neutron has a mass of 1.0087 units. However, when the protons and neutrons make up a nucleus, some of their masses have been converted into energy.

[26]Since isotopes in different concentrations may be present within a sample of an element, any determination of atomic mass may represent an average rather than a specific atomic mass. The mass number of the most common isotope of the element can be usually obtained by rounding off the atomic mass to the nearest whole number.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
The mass of an electron is so small in comparison to the mass of a proton that the mass of the electron is usually disregarded in determining the mass of an atom.[27]	
Electrons are located outside of the nucleus and are distributed in the space which makes up most of the atom.	
In any atom there are as many electrons as protons; therefore, an atom is neutral.	
3 the Bohr atom	
The Bohr model of an atom is only one of a number of atomic models that have been proposed.[28]	*How may the probable structure of an atom be shown according to the Bohr theory?* (H-45)
There are more recent models of atoms which help explain more about the behavior of atoms than the Bohr model can.[29]	
The atomic number indicates the number of protons in the nucleus. The symbol p or + together with a number is often used to indicate the atomic number or the number of protons in the nucleus.	
The atomic number indicates the number of protons in the nucleus. (This is often represented by the symbols, P or +.)	
The difference between the mass number and the atomic number represents the number of neutrons in the nucleus. (This is often represented by the symbol, N.)	
The electrons are located outside of the nucleus. Where they are depends upon how much energy they have.	

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[27] The mass of the electron may be considered to be $\frac{1}{1836}$ of the mass of the proton.

[28] The Bohr model serves quite well to illustrate atomic number, atomic mass, the principal valences, and many chemical combinations of the first 20 elements.

[29] The concepts concerning the Bohr atom refer only to atoms with atomic numbers 1-20. Representations of atoms with atomic numbers greater than 20 are complicated by exceptions to the simple rules stated.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
In the Bohr model electrons are considered to revolve around the nucleus in one of several concentric circular orbits.	
The orbits are called shells or rings and can be denoted by the letters K, L, M, N, O, P, or by the numbers 1, 2, 3, 4, 5, 6.	
Electrons in orbits near the nucleus have less energy than those in orbits farther from the nucleus.[30,31]	
There is a maximum number of electrons which each orbit can hold.	
The first shell (K) can hold no more than two electrons.	
The second shell (L) can hold no more than eight electrons.	
Atoms with atomic numbers 1-20 can have no more than eight electrons in the third shell (M).	
4 classification of elements	
Elements may be classified as metals, nonmetals, or noble gases depending on the number of electrons in the outermost shell.[32]	
Atoms with 1, 2, or 3 electrons in the outermost shell are classified as metals. They will react with nonmetallic atoms.	
Atoms with 5, 6, or 7 electrons in the outermost shell are classified as nonmetals and they will react with metallic atoms.	
Atoms with full outside shells are referred to as noble gases since they seldom enter into any chemical action.	

[30]The quantum wave mechanics model of the atom is beyond the depth and background of most general science pupils. Teachers should be interested in this other model since it gives more useful information than the Bohr atom.

[31]Electrons may be raised to a higher energy level by the addition of energy. When the electron drops back to a lower level, the energy which it has absorbed is released in the form of light. This accounts for different colored flames which are produced when certain substances burn or are heated to incandescence.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
B Chemical action	
The number of electrons in the outermost shell determines the way in which an atom will react chemically.	*How can the probable change in structure during chemical action be shown by a diagram?* (H-46)
During chemical action, the atoms of metals and nonmetals try to form complete outer rings in which condition they are the most stable.	
Metals tend to lend all of the electrons in the last shell in order that a completed outermost shell remains.	
Nonmetals tend to borrow electrons in sufficient quantities to bring the electron number to the maximum amount for the outside ring. The particle gets ring completeness.	
When an atom either borrows or lends electrons, it becomes a charged particle (ion) because the number of electrons no longer equals the number of protons.[33]	
A loss of an electron produces a positively charged ion, while the gain of an electron produces a negatively charged ion.	
There are always two oppositely charged kinds of ions formed during the lending and borrowing of electrons. The two particles arrange themselves to form a new compound.	
C The Periodic Table	
The Periodic Table is used to organize and classify a large body of information about the elements.[34]	*Why was Mosley's work important?*

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[32]Noble gases may also be classified as nonmetals.

Some elements with 3, 4, or 5 electrons in the outermost ring may have properties of both metals and nonmetals.

[33]In the solid state the ions assume definite positions in a crystal lattice. The oppositely charged particles are held in place by electrostatic attraction.

[34]By learning how to use and interpret the Periodic Table, one can avoid memorizing a great amount of detail.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
1 arrangement of the Periodic Table	
The elements in the modern Periodic Table are arranged according to increasing atomic numbers in such a way that similarity is shown between numbers of shells and numbers of electrons in the outermost shell.[35]	*How can pupils understand the arrangement of the Periodic Table?* (H-47)
The elements are arranged in columns and rows. The columns going down the table are called groups. The rows going across the table are called periods.[36]	
Groups contain elements which have similar chemical properties.	
Groups or "families" are often called by the name of the first element in the group; for example group VA is called the Nitrogen Group and VIA the Oxygen Group. Group O contains the noble gases.	*How are members of a group similar in chemical properties?* (H-48)
Some groups have special names such as group VIIA which is known as the Halogens and IA as the Alkali Metals.	
Elements within a group have the same number of electrons in the outermost shell.	
2 use of the Periodic Table	
The location of an element in the table generally indicates the type of element it is.	
Elements found left of the heavy black line, which runs stepwise from boron (B) to the bottom of period 6, are generally classified as metals.[37]	
More elements are classified as metals than as nonmetals.	

[35]Early periodic tables based upon atomic masses led to predictions of properties of elements that were not known at the time. These predictions led to the discovery of many elements.

[36]Sometimes a period is called a series.

[37]Elements found along this heavy black line are classified as metalloids for they exhibit both metallic and nonmetallic properties. These elements act as the "turning point" within groups such as VA and VIA which contain typical metals at the bottom of the group.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Elements between the heavy black line and gorup 0 are classified as nonmetals.	
Elements in group 0 constitute a special group which do not readily react. They are called noble gases.	
Members of group 0 are unique in that they are the only elements to have both electrical balance (equal number of protons and electrons) and a full electron load in the outermost orbit.	
Some atoms have electrical balance but do not have a full electron load in the outside orbit. When these atoms react, they tend to complete the outer orbit in some way.	
The relative activity of an element within a group may be predicted.	*How can the relative activity of an element be predicted?*
Metals usually react because they lend electrons. In those metals whose outer electrons are furthest from the nucleus, the outermost electrons are attracted least by the nucleus and are more easily lost. Therefore, the metals at the bottom of the table are the most active.	
When nonmetals react with metals, the nonmetals gain electrons. In those nonmetals whose outer electrons are nearest the nucleus, additional electrons are attracted to the nucleus most easily. Therefore the nonmetals at the top of the table react most easily with metals.	
IV Common Chemical Changes	
A chemical change produces one or more new substances, each having its own set of properties by which it can be recognized.	

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Two substances may react to form one or more new substances. The products formed depend upon the nature of the reacting substances.	
A Direct union	
Two elements may combine to form a single compound.	*What are some examples of direct union reactions?* (H-49)
A metal combines with a nonmetal to form a compound known as a salt. The metallic atoms lend electrons to the nonmetallic atoms during the reaction.[38]	*What are some common salts formed as a result of the reaction between a metal and a nonmetal?*
Oxygen unites with most elements to form oxides.	*What are silver tarnish and iron rust?*
If a salt is composed of only a metal and a nonmetal, the name of the compound contains two parts. The first part is the name of the metal; the second is the name of the other element with its name ending changed from *-ine* or *-en* to *-ide*.	*Use a model to describe the formation of magnesium sulfide.*
Compounds with names ending with the letters *-ite* or *-ate* contain three or more elements, one of which is oxygen.	
B Single replacement	
In some chemical reactions, one of the elements in a compound is replaced by another element.[39]	*How can a single replacement reaction be shown?* (H-50)
A metal may replace another metal combined in a compound. The products of the reaction are the replaced metal and a new compound.[40]	

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[38] A nonmetal may combine with another nonmetal. Electrons are not loaned but are shared by both kinds of atoms. Carbon dioxide, sulfur dioxide, and the oxides of nitrogen are examples of compounds formed by two nonmetals sharing electrons in covalent bonding.

[39] The reaction between copper and silver nitrate may be expressed as: copper and silver nitrate yields copper nitrate and silver.

[40] There are instances where a metal will combine with another metal. During the formation of an alloy, some metals combine.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
When a nonmetal and a compound containing a metal react, the nonmetal may replace the nonmetal in the compound and a different compound is formed.	
Each element has its own degree of chemical activity.	
An element can replace only those elements which have a lower degree of activity than it has.	*How can the effect of the activity of an element on the rate of a reaction be demonstrated?* (H-51)
An element in a compound can be replaced by any element more active than it is.	
A metal may replace the hydrogen in an acid.[41]	
The elements may be listed according to their degrees of activity. The list is called an activity or electrochemical series. The activity series may be used to predict whether an element and a compound will react.	*What are some pupil activities requiring the use of the activity series to predict whether a reaction will or will not occur?* (H-52)
The most active metals are found at the top of the series.	
A metal can replace any metal listed below it in the activity series.	
The most active nonmetal is found at the bottom of the activity series.	
A nonmetal can replace only those nonmetals listed above it in the electrochemical series.	

[41] Although hydrogen may be classed as a nonmetal, it has many chemical properties similar to those of a metal.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
C Reactions between two compounds	
Two compounds may react to form one or more compounds.	
Usually two compounds exchange parts to produce two new compounds. This type of reaction is called an ionic (exchange) reaction.[42]	*What are some ionic (exchange) reactions?* (H-53)
In order for an ionic (exchange) reaction to occur, a product must be formed that "removes" some of the ions from the solution. This product may be a precipitate, a gas, or water.	
When the solutions of two compounds are combined, one or both of the products formed may be of low solubility.[43]	
A precipitate is a substance of low solubility (insoluble) which may be formed during the reaction between two solutions.	
A solubility table tells the degree of solubility of many substances.	
A substance can be described as soluble, slightly soluble, or insoluble.	
The table of solubilities can be used to predict whether or not a precipitate will be formed during a reaction between two solution.[44]	
An acid and a base reacting to form a salt and water is an example of an ionic (exchange) reaction.	

[42] Pupils can predict the products to be formed by writing a word equation. Have the pupils draw a line between the two parts of each compound's name. The two "ends" combine and the two "middle" ones combine. Remind the pupils that the metal's name is always first in the name of the compound formed.

[43] Actually any substance will dissolve to some extent in water. If, however, only a trace dissolves, it is said to be insoluble.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
D Decomposition	
A single substance may break up to form more than one substance.	*What are some decomposition reactions?* (H-54)
Compounds containing only two elements may decompose to form a metal and a nonmetal.	
Compounds containing more than two elements may decompose to form new compounds or an element and a compound.	
Some substances decompose to form one or more compounds.	
The energy lost during the formation of the compound must be restored during the decomposition process.	
Decomposition reactions usually require energy in the form of light, heat, or electricity.	
V Chemical Shorthand	
Chemical symbols and formulas are used to represent the names of elements and compounds.	
Symbols are used to represent elements and quantities of elements.	
A Symbols	
A chemical symbol stands for the name of a specific element.	
A symbol is a contraction of the name of an element. The symbol may consist of a single capital letter or two letters, the first of which is always a capital letter. The other letter is always a small letter.	*How are chemical symbols used?* (H-55)

[44] After determining the possible products, use the solubility table to see if either product is insoluble. An insoluble product will appear as a precipitate and can be separated from the remaining solution by filtration.

A symbol stands for one atom of an element. A symbol also stands for a gram-atomic mass of an element which is 6.02×10^{23} atoms.

A subscript placed after and slightly below a symbol indicates that there is more than one atom of the element. The number of atoms equals the numerical value of the subscript.[45]

B Formulas

A formula is used to represent the name of a chemical compound.

What information can be obtained from a formula? (H-56)

A formula is composed of symbol(s) and subscript(s).

When a formula consists only of a symbol and a subscript, it represents a molecule of an element. If a formula is composed of two or more symbols together with subscripts, written or understood, it represents the name of a compound.

If a group of symbols is enclosed within brackets or parentheses, each symbol is multiplied by the subscript following the brackets or parentheses.[46]

A formula may represent a quantity of a substance, and stands for one formula-unit of a compound.

C Coefficients

A number preceeding a symbol or a group of symbols is called a coefficient.

A coefficient acts as a multiplier for all of the symbols following it in the compound.

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[45] When a subscript is not written after a symbol, it is understood that the subscript is 1.

[46] The formula $Al(NO_3)_3$, indicates 3 aluminum atoms, 3 nitrogen atoms, and 3 x 3 or 9 oxygen atoms.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
If a formula includes parentheses or brackets, the number of atoms represented by a symbol inside the brackets or parentheses is found by multiplying the coefficient times the subscript following the brackets or parentheses.	
D Chemical equations	
An equation is chemical shorthand for expressing the amounts and kinds of combining substances and the products formed during a chemical action.[47]	*What information can be obtained from an equation?* (H-57)
E Mass relationships	
A symbol may represent one atomic mass unit.	
A symbol with a subscript may represent a mass equal to the atomic mass multiplied by the numerical value of the subscript.	*What is the gram-formula mass for oxygen (O_2) chlorine (Cl_2), and helium (He)?*
A formula may represent a formula-mass unit which is equal to the sum of all of the atomic mass units in the formula.	
In the laboratory, mass is measured in gram-units, not atomic or formula-mass units.	
A symbol may represent a gram-atomic mass.	*Why are gram-atomic masses used in the laboratory?*
The numerical value of the gram-atomic mass is the same as the atomic mass but represents a much larger unit, the gram.	*Which is larger, a gram-atomic mass of aluminum or a gram-atomic mass of magnesium? Explain.*
A formula may represent a gram-formula mass.	
A gram-formula mass has the same numerical value as the formula mass but represents a much larger unit, the gram.	*How many gram-atomic masses are present in a given sample of an element?* (H-58)

[47] For example, 2CO means 2 carbon and 2 oxygen atoms. 2 CO_2 means 2 carbon and 2 x 2 or 4 oxygen atoms. $2HgO \longrightarrow 2\ Hg + O_2$ states that two groups each consisting of 1 mercury atom and 1 oxygen atom forms (decomposes) 2 mercury atoms and 2 oxygen atoms. The different kind of atoms are no longer chemically combined.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Gram-atomic mass and gram-formula masses are used in the laboratory.	*How many gram-formula masses are present in a given sample of a compound?* (H-59)
VI Common Compounds and Mixtures	
A Compounds	
A compound contains two or more elements chemically united.	
A compound has properties unlike its individual component parts; it is a homogeneous substance.	
Elements making up a compound have been united according to a definite proportion by mass.[48]	*How can the combining mass relationship be determined?*
The component parts of a compound can be separated by a chemical change.	
A compound may be an acid, a base, or a salt.	
1 acids	
The properties of a compound determine if it is an acid.[49]	*What are the properties of acids and bases?* (H-60)
Acids turn litmus red; acids neutralize the effect of bases; and, acids combine with many metals.	*How can neutralization be demonstrated?*
Acids play an important part in the manufacture of many products, and are also important in the digestion of our foods.	

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[48] One way of finding the combining ratio uses the correct formula for a compound. Multiply the subscript by the atomic mass of each element. Set the values obtained in ratio form and reduce the ratio to its simplest value. Then set up the proportion, but be sure to keep the order consistent. Example: H_2O
mass of hydrogen = 2 x 1 = 2
mass of oxygen = 1 x 16 = 16
The ratio of 2:16 reduces to 1:8; therefore,

$$\frac{\text{mass of hydrogen}}{\text{mass of oxygen}} = \frac{1}{8}$$

[49] Acids have a sour taste. However, the testing of compounds for this property should not be undertaken by testing. Instead, use an indicator such as pH paper.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
2 bases	
The properties of a compound determine if it is a base.	*What are some properties of bases?*
Bases can turn litmus blue and colorless phenophthalein red; they can counteract or neutralize acids.[50]	
Most bases are insoluble in water.	
Bases may be made by chemical changes.[51]	*How can some bases be prepared?* (H-62)
Bases play an important part in the manufacture of many products.[52]	
3 salts	
A salt contains the positive ion from a base and the negative ion from an acid.	*What are some properties of some salts?* (H-63)
Most salt compounds are found in the minerals of the earth and in sea water.	
Salts may be made by neutralization of an acid by a base.[53,54]	*How can the quantitative treatment of neutralization be shown?* (H-64)
The uses of salts vary from seasoning and preserving food to softening water, making acids or bases, fertilizers, drugs, and explosives.	*What salts are found in most garden fertilizers and plant foods?* *Why is limestone put on soil to make it "sweet?"*[55]
B Mixtures	
A mixture contains two or more elements and/or compounds and may be separated into its constituent parts by physical means.[56]	*Are two mixtures containing the same compounds alike? Explain.*

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[50] A compound such as ethyl alcohol (ethanol C_2H_5OH) containing a covalently bonded hydroxyl group should not be considered a base since it does not affect indicators.

[51] Many bases can be precipitated from their salt solutions by a soluble base. Some metals and metallic oxides can react with water to form bases.

[52] Some bases are used to make soap and neutralize acids. Milk of magnesia is a suspension of magnesium hydroxide and water. It is used to neutralize acids and as a cathartic.

[53] Although salts have a salty taste, most of them are poisonous so the property should not be tested by tasting.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Mixtures exhibiting dual properties are called colloids.	
1 suspensions	
Suspensions are mixtures containing particles which readily settle out of the water (liquid).[57]	*How can suspensions be made and what are their properties?* (H-65)
Suspensions can be separated into their constituent parts by filtering.	
Like all mixtures, a suspension can be made in varying proportions.	
2 solutions	
The particles in a solution are too small to be filtered or to settle out of the liquid.	*What are the properties of a solution?* (H-66)
A solution is clear and uniform; it can be made in varying proportions.	
The solute is the part of a solution that dissolves, while the solvent is the dissolving medium.[58]	
Water is the most common solvent.	
A given amount of solvent can continue to dissolve solute until a point is reached where no more solute can dissolve. Before that point the solution is unsaturated. At that point the solution is saturated at that temperature.	*How can the concentration of solutions be described?* (H-67)
A supersaturated solution contains more solute than it can normally dissolve at the existing temperature. It was prepared by dissolving a solute at a higher temperature and carefully cooling the solution to a lower temperature.	

[54]Normally, commercial quantities of salt are not made in this way. The salts are usually found in nature or produced by other chemical reactions.

[55]"Sour" soil has a pH below 7 and sweetening the soil reduces the acid content.

[56]Some mixtures have properties similar to both solutions and suspensions because the size of the suspended particles lies between those in a suspension.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
The amount of solute that can be dissolved in a definite quantity of solvent depends upon the nature of the substance, the size of the particles, the amount of agitation, the temperature, and the pressure.[59]	*What are some activities in the laboratory which show the factors affecting solubility?* (H-68)
Solutes vary in their ability to be dissolved in a solvent.[60]	
If other factors are the same, small particles of a solute dissolve more rapidly than large particles.[61]	
A solubility curve (graph) is a mathematical picture which shows that a definite amount of solvent can dissolve different amounts of solute at various temperatures. The solution in each case is always saturated since no more solute can be dissolved at the given temperature.[62]	*How can pupils be introduced to the use of solubility curves?* (H-69)
The horizontal axis of the graph represents (unless stated otherwise) the centigrade temperature of the solvent.[63]	
The vertical axis of the graph represents (unless stated otherwise) the grams of solute that can dissolve in a specific amount of solvent, usually water.	
The data used to make solubility curves have been obtained by experimentation.	*How can solubility be plotted?*[64] (H-70)
Solubility curves are used to predict the amount of solute that may dissolve in, or precipitate out of, solution as the temperature is changed.	*How can solubility curves be used in solving problems?*[65] (H-71)
A solubility curve shows the temperature at which a given amount of solute will saturate 100 grams of water.	

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[57] True suspensions are opaque to the passage of light and their particles are visible. If during filtration of a suspension the filtrate appears to be cloudy, some suspended particles are small enough to pass through the filter and the filtrate is really a colloidal dispersion. The original mixture was, therefore, both a suspension and colloid.

[58] The solute may be in the gaseous, liquid, or solid state. Likewise, the solvent may be in any one of the three states. Therefore, there are nine different types of fundamental solutions. All pupils at this grade level should consider the solutions in which the solvent is a liquid.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
A solubility curve can be used to show how much additional solute is needed to keep a solution saturated as the temperature increases.	
A solubility curve can be used to show how much solid will crystallize out as a saturated solution cools.	
3 colloids	
Some mixtures have properties similar to both solutions and suspensions because the size of the particles lies between those in a solution and those in a suspension. These mixtures are called colloids.	*What are the properties of colloids?* (H-72)
The path of a beam of light can be seen as the light passes through a colloid.	
A mixture may be a combination of a colloid and a suspension, a solution and a colloid, a solution and a suspension, or a suspension, solution, and a colloid.	
VII Introduction to Organic Chemistry	
Organic chemistry is the chemistry of the compounds of carbon.	
A Organic compounds	
Although all living things make organic compounds, most organic compounds used by man are synthetic.	*How can organic substances be recognized?* (H-73)
There are many more organic compounds than inorganic compounds.	
Virtually all new chemical compounds being discovered can be classified as organic compounds.	

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[59] Solids usually become more soluble in a given amount of solvent as the temperature increases. Gases become more soluble if the temperature of the solvent decreases providing the pressure remains the same. Increasing the temperature increases the ability of the solute to dissolve since molecular action at the surface of the crystal is increased as the temperature increases.

[60] Polar solvents such as water can dissolve polar solutes such as salts of sodium and potassium.

Nonpolar, organic chemicals such as grease dissolve best in nonpolar solvents. Nonpolar solvents include chloroform and kerosene. Alcohol is a molecule, one end of which is polar, the other nonpolar. Alcohol dissolves many polar and nonpolar substances.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES

B Some differences between organic and inorganic compounds

There are many more (about 20 times more) organic compounds than inorganic compounds.

How can differences between organic and inorganic compounds be illustrated? (H-74)

Organic compounds are more easily decomposed by heat than inorganic compounds.

Organic compounds char when heated.

Solutions of organic compounds are usually poor conductors of electricity.

Organic compounds are usually only slightly soluble in water.

Organic compounds tend to react more slowly than do inorganic compounds. A notable exception is the speed with which some organic reactions occur in living things.

C Bonding in organic compounds

1 covalent bonds

Some atoms tend to combine by sharing electrons A pair of electrons shared between two atoms constitutes a covalent bond between the atoms.

How can the differences between ionic and covalent bonds be shown? (H-75)

Atoms tend to share electrons in sufficient quantity to bring the electron number to the maximum amount permissible for the outside shell of each atom bonded.

The maximum number of electrons that the outer shell of the hydrogen atoms can hold in combination is 2; for all other atoms the number is 8.

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[61] Pulverizing the solute increases the surface area that can be exposed to the solvent.

[62] The amount of solvent may be expressed in 100 ml or 100 grams of solvent.

[63] Several curves may be plotted using the same set of axis. However, the name of the solute must be written on each curve since each solute has its own individual solubility curve.

[64] The data for the solubility reference curves in appendix (H-70) may be used by pupils in drawing their own solubility curves.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
Organic compounds and some inorganic compounds form by covalent bonding.	
2 formation of molecules	
Some molecules are formed by covalent bonding.	
a hydrogen molecule	
Two hydrogen atoms join by a covalent bond to produce a molecule.[67]	*What are some ways to represent the structure of the hydrogen molecule?*
A pair of shared electrons may be represented by a dash (—) or by two dots (:).	
A structural formula represents the atoms present in a molecule and the bonds joining them.	
b methane molecule	
The carbon atom has four electrons in its outer shell. Carbon tends to form compounds with other atoms by four covalent bonds.[68]	*How is the structure of a methane molecule represented by molecular models and by a structural formula?*
In methane (CH_4) the carbon atom is joined to each of four atoms of hydrogen by a covalent bond.	
c formation of long chains	
Carbon atoms have the ability to produce long chains or rings by forming covalent bonds with other carbon atoms.[69]	
The ability of carbon to form chains or rings accounts for the number of possible compounds.	

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[65] In performing these exercises, pupils may use a simple calculation. Multiply the mass of solute for 100 gm of water by the mass of water desired and divide by 100. Another way is to set up a straight line graph. Label the amount of solvent on the vertical axis and the mass of solute on the horizontal axis. Plot the point expressing a given solvent-solute relationship on the solubility curve. Connect the origin and the point and extend the line. To use the new graph, read across the line representing the mass of solvent to be used until the straight curve is reached; read down to the horizontal axis and find the mass of solute needed.

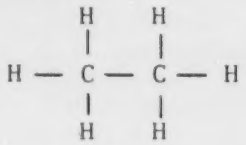
CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES

d saturated and unsaturated compounds

Each carbon atom in a compound may be joined to another carbon atom by sharing one pair of electrons (one covalent bond). This compound is considered saturated.

What are some compounds that are saturated?

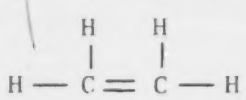
An example of a saturated compound is ethane (C_2H_6) represented as:



Note that there is only one bond between the carbon atoms.

Two carbon atoms in a compound may be joined by sharing two or three pairs of electrons. This compound is considered unsaturated.

Examples of unsaturated compounds are C_2H_4 and C_2H_2. Their structural formulas are:



and

Note the double and triple bonds between the carbon atoms.

D Some classes of organic compounds

Classes of organic compounds may be recognized by the number and arrangement of certain atoms or groups of atoms in their structural formulas.

---

[66]A large percentage of the time devoted to this topic should be with molecular models and visual devices.

[67]Other gases with molecules having covalent bonding include oxygen, nitrogen, fluorine, and chlorine.

[68]A satisfactory device to illustrate the arrangement of covalent bonds in a variety of organic compounds is a set of molecular models. The models consist of wooden spheres with one to five holes in them and dowels or springs to represent the covalent bonding. Sets may also be constructed of styrofoam spheres and dowels.

[69]This can be visualized by hooking 20 to 40 paper clips together. Each clip represents a carbon atom.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
The properties of organic compounds are related to characteristic functional groups in their molecules.	
1 hydrocarbons	
Hydrocarbons are compounds of hydrogen and carbon.	
Hydrocarbons may be in the form of straight chains or rings, and may contain from one to over 40 carbon atoms.	
Other carbon compounds may be considered as derivatives of the hydrocarbons, even though they may not be manufactured from them.	
There are many hydrocarbons which are unsaturated.	*What are some hydrocarbons that are unsaturated?*
2 alcohols	
Alcohol may be formed by oxidizing hydrocarbons.	
Alcohols may be derived from saturated hydrocarbons by substituting one or more — OH groups for one or more hydrogen atoms.	
Alcohols may result from the addition of OH groups to unsaturated hydrocarbons.	*Show by diagram or molecular models how one H atom in ethane (C_2H_6) is replaced by an — OH to produce ethyl alcohol (ethanol).*
The general structural formula of an alcohol is R — OH where R is any hydrocarbon minus one H.	
a methyl alcohol	
The formula for methyl alcohol (methanol) is CH_3OH.	

Methyl alcohol is used to make other organic compounds, and as a solvent and automobile antifreeze. It is very poisonous.

b ethyl alcohol

The formula for ethyl alcohol (ethanol) is C_2H_5OH.

Ethyl alcohol is used to make other organic compounds and is a solvent.

Ethyl alcohol is a narcotic.[70]

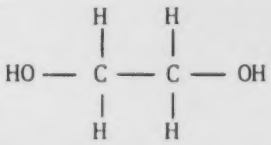
How does ethyl alcohol affect the human body?

c other alcohols

Ethylene glycol, $C_2H_4(OH)_2$, and glycerol, $C_3H_5(OH)_3$ are examples of other alcohols.

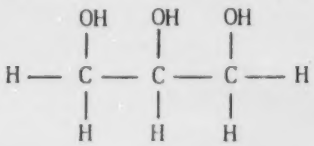
Ethylene glycol is used as an automobile antifreeze.

The structural formula of ethylene glycol is:



This compound has 2 — OH groups.

The structural formula of glycerol is:



This compound has 3 — OH groups. It is a part of all body fats.

---

[70] Reference should be made to the adverse effects of alcohol on the human body. Some of these adverse effects should be explored in detail.

CONCEPTS AND UNDERSTANDINGS

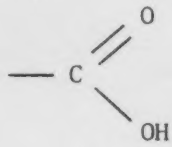
QUESTIONS AND ACTIVITIES

3 organic acids

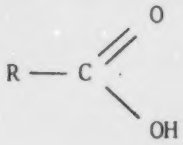
Organic acids may be prepared by oxidizing alcohols.[71]

What are some common organic acids?

Organic acids have a characteristic carboxyl group:



The general formula for an organic acid is:

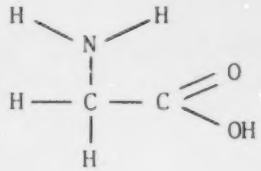


Some organic acids are the fatty acids which are the building blocks of fats.

a amino acids

Amino acids are organic compounds in which an amino ($-NH_2$) group has replaced a hydrogen atom.

The amino acid derived from acetic (ethanoic) acid is represented by the formula:



Amino acids are units which make up proteins.

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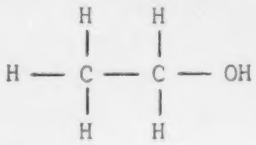
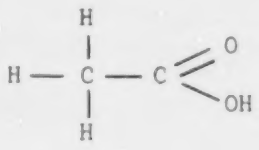
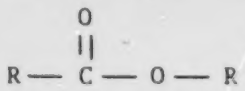
[71] Although not all acids are produced directly from the oxidation of an alcohol, the general process can be represented as follows:
Start with ethyl alcohol,



If oxygen is added, one atom of oxygen joins with two atoms of hydrogen to produce water and acetic (ethanoic) acid,



CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
b esters	
Esters can be formed by the reaction of an organic acid and an alcohol.	*How can esters be prepared in the laboratory?* (H-76)
The general formula for an ester is: 	
Animal and vegetable fats and oils are esters which are made up of glycerol and fatty acids.	
VIII The Role of Chemistry in Society	
A Use of chemicals	
Every living organism depends upon chemical processes.	
B Conservation of natural resources	
Most of our natural resources are being rapidly consumed (or made difficult to reuse) as the demand for chemical products increases.	*Which natural resources are being used at a very fast rate? Why must we be concerned about their conservation?*
Scientists are constantly searching for ways to conserve our stockpile of resources.	*What new careers are developing in the area of chemical resource conservation?*
Recycling is an important means of conserving chemical resources.	*What are some approaches being used to recycle chemicals?*

Some common acids are formic acid (HCOOH), found in bees and ants, and acetic acid (CH_3COOH), found in vinegar.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
1 using substitutes	
Using substitutes conserves a resource in short supply.	*Consider the use of substitutes or alternatives in the context of recurring crises pertaining to fuel resources and mineral resources.*
2 using protective coatings	
Protective coatings prevent or slow down chemical reactions which reduce the usefulness of a substance.[72]	*Discuss the relative value of petroleum when used as a fuel resource and when used in the petrochemical industry. Describe this relationship in terms of the need to develop alternative energy sources.* *Why do rust spots appear on cars? What causes the rust? How is rusting overcome?* *Explain how economic objectives (ever-increasing automobile sales) sometimes thwart the efficient use of technological advance (effective rustproofing processes).*
3 salvaging waste products	
The effective use of waste products reduces the drain on our resources.	*What products are being recycled in your community? What could be recycled?*
4 proper disposal of waste products	
Proper disposal of waste products prevents pollution of our natural resources in the air, water, and the earth.	*What is a sanitary landfill? Why are only biodegradable detergents used today?*

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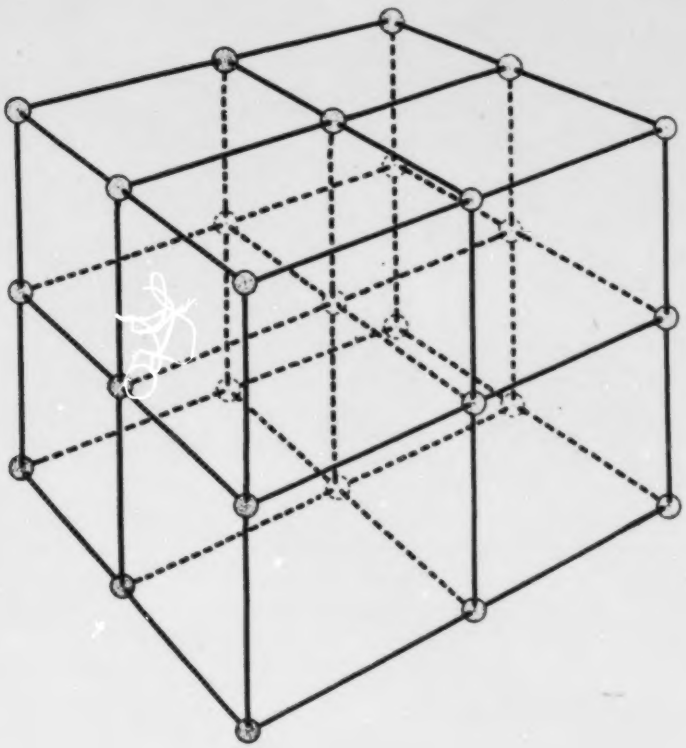
[72] The corrosion of metal is a serious problem. Since most pupils are interested in automobiles, the corrosion of metal on automobiles and ways to prevent the corrosion provide a motivating approach to the topic of conserving metals.

There are several films on the corrosion, conservation, and refining of metals available from film libraries.

CONCEPTS AND UNDERSTANDINGS	QUESTIONS AND ACTIVITIES
	Why are we concerned about water pollution? What are the sources of air and water pollution?
C The balance of nature	
The use of chemicals to improve some feature of man's environment may affect seriously some other feature.	*Why have certain insecticides and all phosphate detergents been removed from sale in New York State?*

APPENDIX - LEARNING ACTIVITIES



H-1. Prevention of Accidents in the Laboratory [2, 8, 9]

Ask the students to describe some accidents that have occurred in their homes, school, or community. Then have the class discuss the possible cause(s) of each accident and decide whether or not each accident might have been prevented. Develop a set of generalizations which are based upon their answers. Expand the generalizations into some laboratory safety rules. Some generalizations which might be included are:

- Be aware that your actions may affect others.
- Inattention to the task at hand may cause an accident.
- Lack of knowledge may cause an accident.
- Working too fast or too slow may cause an accident.

H-2. Preventing Chemicals from Entering the Body [1, 2, 14]

Some laboratory safety rules which apply to the prevention of chemicals entering the body may be developed from the following activities:

- Observe the odor of the hands before and after peeling and chopping an onion. Hold one hand under running water for 5 to 10 minutes. (Ask the class why both hands are not held under the water.) Compare the odors of each hand. (The need for safety glasses might be discussed since the onions may cause the eyes to smart. Point out that many chemicals can burn or irritate the eyes thus, safety glasses are worn in a laboratory.)

- Spread a few drops of water-soluble ink over a piece of notebook paper. When the ink has dried, cut the paper into two strips. Let running water flow over one strip for a few minutes. Compare the amount of ink on each strip. Discuss why the original paper was cut into two strips.

- Show that vapors can be adsorbed on food: Into each of two similar jars, place some pieces of bread. Cover one jar. Into the second jar place a drop of perfume or oil of wintergreen so that the bread does not touch the liquid. Cover the jar. After 5 to 10 minutes, remove the bread from each jar and compare them. Account for any difference in odors. Remind pupils that many of the chemicals that they will use can release toxic vapors into the air and could be adsorbed on any food open to the air. (Chopped onions may be substituted for the perfume or oil of wintergreen.) The need for the control to be exactly the same except for one variable might be considered. Would different amounts of bread in the jars give different results? Use much more bread in the jar with the vapor source than in the control jar. It might be advisable to use a less pronounced odor source as well.

- Before the class begins, set up 3 to 5 beakers, each containing a different, colorless liquid. The series of beakers might contain dilute solutions of each of the following: hydrochloric or

sulfuric acid, sodium or potassium hydroxide, sodium carbonate, aluminum chloride or sulfate. One beaker must contain only distilled water.

- Ask the students to point out which beaker(s) contains only distilled water. Whatever their hypothesis, it may be tested by the addition of the same number of drops of universal indicator. If the pupils are unacquainted with this indicator, discuss what information must be gained before the indicator test is valid. (Observe the color of the indicator in distilled water for a comparison or control.)

- After the class has identified the distilled water, let the pupils read the labels on the containers of solutes to find out the poisonous nature of the chemicals. Labels on containers should give this information.

- Wash some glassware with soap but do not thoroughly rinse away the soap. Pour 1 or 2 ml of phenolphthalene into the glassware. Explain why a color change occurs. Relate this activity also to the need for using clean glassware to avoid contamination of the chemicals to be used.

The above activities should lead to the development of the laboratory safety rules:

- Do not eat or drink anything in the laboratory.
- Flush any chemical spilled on the skin with quantities of water.
- Wear safety goggles.
- Use clean glassware.

H-3. Effect of Chemicals on Clothing [1, 2]

Give each pupil small pieces of cloth such as nylon, cotton, linen, silk, wool, and a blended fabric. Assign each pupil a different chemical to drop on each sample of fabric. Observations should be made at once, after a period of 1 to 2 days, and also after the fabrics have been washed. The weakening of some fibers may be observed by comparing the passage of light through the treated and nontreated areas. Chlorox or other chlorine bleach, sodium hydroxide, and acids affect many fabrics. Yarn and discarded hose may be substituted for some of the above fabrics.

H-4. Contamination by Chemicals [9, 16, 17]

Before the class period, place a few drops of phenolphthalene indicator on the glass stoppers of some reagent bottles. After the stoppers have become dry, put them back into the bottles. During class, have some pupils prepare a basic solution such as sodium carbonate and pour the solution into the bottles for storage. Ask the pupils to account for any changes that they observe as the liquid is poured into the bottles.

The effectiveness of this problem to be presented to the class depends upon the preparation and testing before class. Into each beaker place a drop of the solution indicated on the following page.

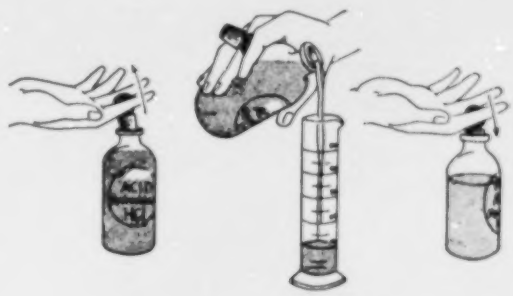
Beaker #1	1 drop 0.001 Molar phosphoric acid
Beaker #2	1 drop 0.01 Molar hydrochloric acid
Beaker #3	3 drops 1 Molar sodium hydroxide

The demonstration consists of pouring distilled water, to which two or three drops of universal indicator have been added, into beaker #1. Then from beaker #1 into beaker #2 and then into beaker #3. It may be necessary to adjust the drops of reagents to get the distinct color changes; hence, the need for pretesting because the molarity of reagents and the amount of dissolved gases present in the distilled water may affect the results.

The class should be able to make observations; to make hypotheses which explain why there were changes as the water was poured into apparently clean glassware; and to identify what safety rule was involved in the activity.

H-5. Pouring Liquids [4]

Demonstrate the proper procedure to use when pouring liquids from a bottle. Then have the pupils measure specific amounts of the liquids from the reagent bottles.



H-6. Use of Safety Devices [4]

Point out where the safety devices such as a fire blanket, fire extinguisher, and safety shower are located and instruct the pupils in their use. Point out that a carbon dioxide fire extinguisher is never directed at a person because serious cold burns can result from the solid dry ice touching the skin. When a person's hair or clothing is on fire, apply the fire blanket from the head downward to prevent fanning flames toward the hair. Ask pupils how to report a fire in the laboratory and have them follow the fire routes to be used. Stress the need to keep calm and to listen for changes in direction. The local fire department may provide demonstrations on putting out fires with fire extinguishers.

H-7. Reporting Accidents [8]

Instruct pupils in the procedures that are to be followed in case of an accident. Show them the accident report form and discuss how it should be completed. Stress that the knowledge gained about an accident can help prevent a similar accident from occurring. A suggested sample form is shown on the following page.

Centerville Central School
Centerville, New York

Accident Report Form

Name of Pupil_____

Age_____Time_____Date_____

Describe accident and where it happened_____

_____

_____

Describe apparent extent of injury_____

_____

_____

Names and addresses of witnesses_____

_____

_____

What first aid was applied_____

_____

What instructions were given to the pupil_____

_____

Pupil transported by_____To_____

Teacher in charge of pupil at time of accident

_____

H-8. Comparison of Heat Output [1, 2, 5]

The rate of heating (degrees centigrade per second) varies directly with the intensity of the heat source.

Set a 250 ml beaker or tin can containing 100 ml of water on an iron ring stand adjusted so the ring is 8 cm above the heat source. Suspend a thermometer so the bulb is in the center section of the water. Use a burner as the heat source for one setup and a candle for a second setup.

Divide the class into two teams, each being responsible for the observations of one setup. Each team records the initial temperature of the water and the temperature every 20 seconds until the water in one setup is 90°C. The rate of heating may be calculated:

$$\frac{\text{Final temperature} - \text{initial temperature}}{\text{total time of heating}} = \frac{\text{°C gain}}{\text{seconds}}$$

The two sets of data may also be graphed on the same set of axes with time on the horizontal axis. The slopes can be used for a rate of heating comparison. Some of the more capable students may consider sources of experimental error such as heat loss and low conductivity of heat by the glass container.

H-9. The Burner and its Flame [2, 4, 6]

Allow each pupil to examine, dismantle, and reassemble a burner to identify its parts. After having observed the proper way to light and adjust a burner, each pupil should ignite and adjust the burner several times.

Have the pupils study the different flames produced as the amount of air is varied.

Place a Pyrex beaker of cold water on a ring stand. Direct the blue flame toward the outside of the beaker. Gradually decrease the air intake. The flame will deposit carbon only when the flame is yellow. In this case, there is not enough air intake to completely burn the fuel.

Place a wooden splint across the barrel of a burner which is producing a blue flame. As the splint starts to scorch, remove it from the flame, and study the pattern left on the splint. Compare the result with scorch patterns produced when the splint is held at different levels in the flame. Ask the class to propose a reason for the scorch patterns produced. If the splint starts to burn, quickly remove the splint from the burner flame and extinguish the fire. The splint will always catch on fire at the tip of the blue cone. Ask where the hottest part of the flame seems to be.

H-10. Flame of a Candle [9, 14, 16]

Have the pupils examine through a magnifying glass the flame of a candle. Using tongs, hold a slightly dampened white card horizontally in the flame so that one end of the card touches the wax. Remove the card as soon as it starts to char. Compare the char pattern with that which a burner flame produces on a wooden splint. Identify the hottest part of the candle flame.

Ask the class to develop a model "flame" and to compare the model with the flame of a burner.

Ask the pupils to predict what the flame of a lighter looks like, then have the class look at the flame through a magnifying glass. The model produced by the pupils may be compared and revised if necessary by studying other flames. (Require safety goggles and observe fire precautions.)

H-11. Heating Liquids in a Test Tube [1, 4]

Demonstrate the need for observing the rule: "Never point a test tube being heated toward a person." Pour some water in a test tube. Rapidly heat the water so it splashes out of the tube onto the demonstration table.

Demonstrating the "wrong way" is often an impressive teaching method and emphasizes the need for caution.

H-12. Affect of Heat on a Reaction [5, 12, 13, 14]

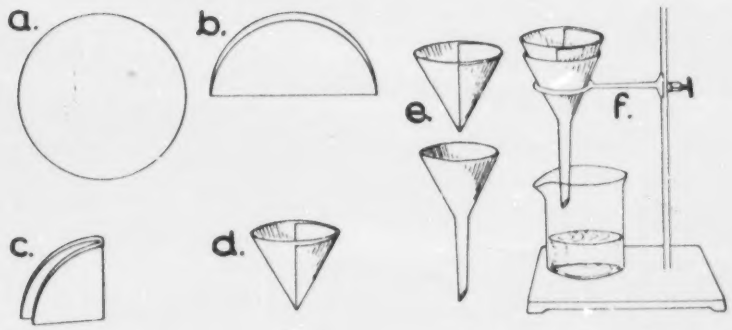
Pour 200 ml of cold water (10°C or less) into a liter beaker and 200 ml of hot water (80°C-100°C) into a second beaker. Record the temperature of each sample of water. Simultaneously add 1 tablespoon of baking powder to each beaker. Do not stir. Compare the levels of foam produced in each beaker. Ask the class to account for the difference in the heights of the foam and the apparent difference in reaction rate. Relate the activity to appropriate safety rules.

This activity may be extended by using a series of beakers in which the water is at different temperatures. Graph the heights of the foam (vertical axis) vs. the temperature (horizontal axis).

Test the generalizations made by studying the rate at which colloidal sulfur is precipitated. On a white sheet of paper draw a heavy line and protect the paper with a piece of clear plastic. In a 100 ml beaker place 20 ml of hypo solution (20 grams per liter of known temperature), add 1 ml of 6 Molar hydrochloric acid, and stir. Start timing the reaction as soon as the acid is added. Consider the reaction completed when the line is no longer visible. Repeat, using hypo that has been stored in an ice bath to obtain a low temperature. Repeat with a third solution which has been warmed by a hot water bath to about 40°C. Graph the results with time on the horizontal axis. After the curve has been made ask the class to predict the reaction time for the hypo solution at a temperature of 15°C. Have pupils experimentally determine whether their prediction was right. Extrapolate the curve and predict the reaction time at 60°C. (Sulfur dioxide is released so the area should be well-ventilated. Discard solution into the sink and rinse it down the drain as soon as the reaction stops.)

H-13. Use of a Funnel [3, 4, 6, 11]

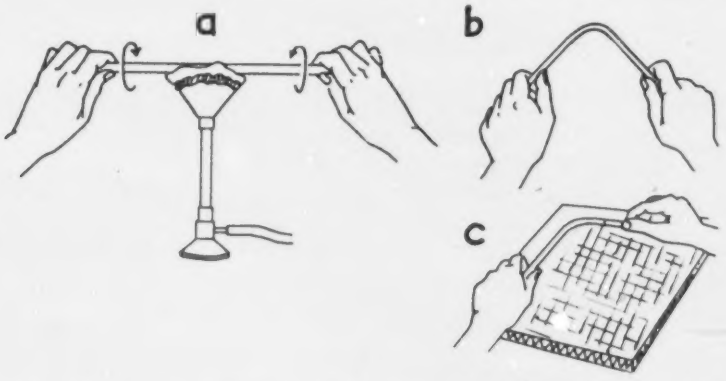
Set up a filter to separate the precipitate from the solution obtained in the previous activity. Demonstrate the method of folding and using filter paper. A piece of paper towel can be used instead of filter paper. If possible, all pupils should have the opportunity to carry out this activity. A sequence of procedures is shown on the following page.



Have pupils filter a number of samples to determine when filtering is effective. Examples may include muddy water, rain water, a colored solution, and a chalk-water mixture, a precipitate formed by adding sodium carbonate to a calcium chloride solution.

H-14. Working with Glass Tubing [4]

Demonstrate each technique before letting the class perform the operation.



(a) Cutting Glass Tubing. At the desired distance from one end of a length of glass tubing, make a deep scratch by drawing the file across the glass as it it held firmly on a flat surface. Hold the tubing so that the cut is away from the body. Grasp the glass so that the thumbs are behind the scratch and the fingers are extended over the glass. Using the thumbs as a fulcrum, gently pull with the fingers toward the body. The glass should break easily at the scratch mark.

If the scratch is not deep enough, a clean break will not be made. The glass tubing may be held in a towel to avoid injury to the hands and finger.

After breaking the glass tubing, immediately fire polish the cut end.

(b) Fire Polishing Glass. The ends of glass tubing should always be fire polished.

Hold the tubing so that one end is just above the blue cone of a burner. Rotate the glass, and remove it from the flame as soon as the end appears to be smooth. Too much heating may result in the tube becoming sealed.

Caution pupils about touching hot glass as serious burns may occur. Do not allow more than one pupil to heat glass in a burner at the same time.

(c) Bending Glass. Fire polish the ends of a 20-centimeter length of glass tubing. Now with a wing top on the burner, hold the glass tubing lengthwise above the blue cone as shown in diagram (a). Rotate the tubing with both hands, and heat the center section until it becomes soft and pliable. Remove the glass from the flame and bend it to form a right angle. If the hot glass is placed on a wire gauze, the corner of the square can be used as a guide for making the bend. Discard any bend in which the glass bore has been pinched.

Emphasize the need for rotating the glass with both hands while it is being heated. Demonstrate what happens when the glass is held firmly in one hand and rotated with the other hand during the heating. The glass will coil up and cannot be bent.

Compare glass bends made by using the blue cone of a burner with those made with a flame coming from the wing top. Note the bore of the tubing in the bent part of each right angle bend. Ask the pupils to explain why the wing top bend is preferable.

H-15. Evaporation of a Solution [4]

Have the pupils assemble the apparatus needed for evaporating a solution. Ask them to recall the safety rules to be observed during the use of a burner and emphasize the need to keep the face away from the work area since there is always the danger of the spattering of a hot solid during the evaporation process.

About 3 ml of sodium chloride or copper sulfate solutions can be evaporated on a watch glass over a beaker of boiling water. Remind pupils to remove the heat source before all of the liquid has been evaporated to prevent possible decomposition of the solid.

H-16. Using and Storing Household Chemicals Safely [9, 13, 16]

Display containers of a chlorine bleach, a sodium sulfate (or bisulfite) toilet bowl cleaner such as solid Vanish, an insecticide, and a latex-based paint. Ask the class whether each item is safe to use, then have pupils read the labels on each container and decide whether it is always safe to use each substance. If not safe always, ask the class to describe the conditions needed for safe usage.

Have the pupils make an inventory of all household cleaners, detergents, bleaches, insecticides, aerosols, waxes and polishes, foams, sprays, drain cleaners, cold tablets, and capsules, vitamins, aspirin, laxatives, cough syrups, and nose drops found in their homes. The inventory may be recorded on a form which will insure that the pupils' read the labels.

		Information from Label		
Substance	Where Stored at Home	Recommended (or not) Storage Area	Special Precautions	When Harmful

As the pupils report their findings, complete a master chart, and then ask the class to decide whether any are unsafely stored. Consider temperature, availability of an open flame such as a pilot light on a stove, ventilation, flammability of the substance, toxicity, and accessibility for creeping babies, toddlers, and pets. Remind the class that many accidents occur to small children who pull household chemicals from under sink storage areas and who ingest the substances or spill them on the skin or into the eyes with frequently fatal results.

An alternate approach is to give the class a list of items commonly found in a home and ask them to write down where they would store each item at home. Have students read the labels on containers of each substance and decide whether they have selected a safe storage area.

H-17. Exhibiting Matter [13]

Make an exhibit of the different kinds of matter. Such a display might include the following:

Elements	Compounds	Mixtures
copper	water	copper sulfate and marble chips
sulfur	copper sulfate	salt and sand or sawdust
aluminum	table salt	muddy water
iron	sand	iron filings and chunks of roll sulfur
mercury	table sugar	pieces of charcoal and sugar
tin	marble chips	
zinc		

H-18. Comparing Elements and Their Compounds [13]

Prepare a display composed of a bottle or jar of each of copper, sulfur, and copper sulfate. Point out that the elements cannot be recognized in the copper sulfate. Other compounds and their component elements can be displayed in a similar manner.

H-19. Endothermic Phase Change [1, 2, 5, 9]

Have a pupil hold an ice cube. As the ice melts, ask him to explain why his hand is cold.

Put some crushed ice into a small, damp paper bag and place it on a wire gauze on an iron ring stand. Gently heat the bag. Ask the class to propose

a reason why the bag doesn't burn. Continue heating the bag until the water boils. The bag will burn after the water has boiled away.

Have the students put some rubbing alcohol on their skin and explain why their skin becomes cold.

Insert a thermometer into a one-hole stopper in an unglazed porcelain cup containing water. Pupils record the initial temperature of the water, then set the apparatus in the air flow from a fan. Have the pupils propose an explanation for any temperature change in the water.

TEACHER DEMONSTRATION ONLY. The following must be done under a well-ventilated hood. On a wooden block, place 2 or 3 drops of water. Set a watch glass on the water on the block then pour 2 to 3 drops of ether on the watch glass. After the ether has evaporated, inspect the water under the watch glass. Ask the class to explain why the water froze and what phase changes had occurred.

H-20. Comparison of a Solid and a Liquid [5, 4, 7]

Have the pupils inspect a rectangular block of wood, then measure the dimensions of the block of wood and calculate its volume. Determine the mass of the wood. Place the wood into various shaped containers to see if it changes its shape or volume.

Give the students a sample of water in a rectangular shaped container of known inside width and length and mass. After determining the mass of water in the container, have the pupils measure the height of the water and calculate the volume of water in cm^3 units which can then be equated to ml units. Carefully pour the water from the container into a graduated cylinder of appropriate size and known mass to see whether the liquid keeps the same shape. Determine the mass of the water to see if any mass is lost. If so, ask pupils to account for the loss.

H-21. Melting Points of Solids [4, 5, 7, 14]

Stir boiling water in a small beaker with a strip of zinc or copper, then stir the boiling water with a stick of Wood's metal.

Place a different metal into each of three crucibles on separate iron ring stands. Metals used should be copper, tin, and lead. Start heating each metal at the same time and notice which melts first.

Have pupils inspect samples of sodium chloride, water, and a gas such as carbon dioxide. Ask the pupils which have melting points higher than room temperature, which has the lowest melting point.

Compare the melting points of a corn oil margarine and butter. (Put the fats in a refrigerator for a few hours before classtime.) Put each cold, solid fat in a test tube in a water bath. Insert a suspended thermometer into each fat. Gently heat the water and observe which fat starts to melt first. (Melting points will not be sharp since the fats are not pure substances.)

H-22. Temperature and Changes of State [2, 12, 13, 14]

(a) Have pupils place some crushed ice into a 150 ml beaker. Support a thermometer so that its bulb is buried in the crushed ice. At 1 minute intervals have the pupils read and record the temperature and note the phase(s) of matter seen in the beaker at the time of each reading. Continue to take the temperature after the ice has disappeared.

The data may be graphed with temperature on the y-axis and the time on the x-axis. The graph will be a horizontal line during the change of state since any heat absorbed by the ice is used to rearrange the molecules for the liquid state. After all the ice has melted, the temperature will rise at a fairly uniform rate. Details of the latent heat of fusion can be found in physics textbooks.

The time required for this activity can be decreased by setting the beaker of ice in a dish of hot water.

(b) TEACHER DEMONSTRATION ONLY. Support a 500 ml Pyrex Florence flask on a ringstand. Pour about 300 ml of water into the flask. From a ringstand clamp above the flask hang two thermometers so that the bulb of one is in the water and the bulb of the other is up in the neck of the flask. Heat the unstoppered flask and its contents. As the water boils, compare the reading of the two thermometers over a period of 2 or 3 minutes.

Under no circumstances put the thermometers into a stopper inserted in the flask. Pressure will build up and the flask would probably "blow up." Remember that steam can cause serious burns. Avoid getting any part of the body near the steam escaping from the flask.

H-23. Boiling Points of Liquids [5, 6, 7, 9]

Under a hood, set up a distilling apparatus with a water condenser. Insert a thermometer through the one-hole stopper in the neck of the flask so that the thermometer bulb is by the delivery tube part of the distilling flask. Pour some isopropyl alcohol (propanol-2) into the distilling flask and heat the flask until vapors just begin to condense in the water condenser. Record the boiling temperature. Repeat the operation using water and compare its boiling temperature with that of the alcohol.

H-24. Boiling and Freezing Points [14, 15]

Point out to pupils that there are many places in which the data for boiling and freezing points of different substances can be located. Show them some samples of charts found in the high school chemistry and physics textbooks.

Have the pupils look up the boiling and melting points for substances like bromine, lead, tin, sodium chloride, carbon dioxide, sand, carbon tetrachloride, and ethyl alcohol (ethanol). Ask the pupils to predict the phase of each substance at: room temperature (20°C), 200°C and -30°C.

H-25. Physical Properties of Some Metals [1, 4, 6]

(a) Tensile Strength of Metals. Illustrate the differences in tensile strength between different metals. Use two wires of the same length and diameter (No. 24 or 28 gauge). Support the wires from one end. Add weights until each breaks. The tensile strength varies directly with the force needed to break the wire, and is a measure of the cohesive force between adjacent molecules over the entire cross-sectional area. CAUTION: Be sure to wear safety goggles.

(b) Malleability of Copper. Demonstrate the malleability of copper. With a ball peen hammer pound a sheet of light-gauge copper into an ashtray. The wooden form for ashtrays is usually available from the art department or school shop. CAUTION: Be sure to wear safety goggles.

(c) Malleability of Lead. With a ball peen hammer pound a lead fishing sinker into a "sheet."

(d) Ductility of Metals. Illustrate ductility by passing around wires of different gauges and composition.

(e) Changing Ductility by Heating. Bend a bobby pin back and forth several times to show its ductility. It is difficult to break it by bending and flexing. Heat the pin to redness in the bunsen flame and quench by plunging the pin into cold water. The bobby pin is brittle now and can be broken easily into pieces.

(f) Testing Hardness. Ask pupils to look up in an earth science or geology textbook the hardness scale used to test minerals. Then, using the hardness standards, determine the hardness of several materials, including some metals like copper, iron, and aluminum.

If the hardness scale testing outfit is not available, some hardness ratings can be roughly obtained by the following:

Matter is scratched by a fingernail	-	Hardness rating 1 - 2
Matter is scratched by a copper penny	-	Hardness rating 3 - 4
Matter is scratched by a steel knife	-	Hardness rating 5
Matter scratches glass	-	Hardness rating over 6

H-26. Detecting the Presence of Iron in Oil [2, 9, 14, 17]

A mixture of oil and fine iron filings is placed into two small aluminum dishes. Holding one dish so the bottom touches the poles of a strong U magnet, gently swirl the oil around. While the magnet is still in place, punch a hole in the side of the dish so the oil drains out into a dish. Repeat the procedure without the use of the magnet. Compare the residues left in the dishes.

Compare the above results with those obtained when the oil-iron mixture is swirled in a "tin" can with and without the use of a magnet. Relate the activity to the way industry can detect the presence of magnetic impurities in fluids and also tell whether iron and steel machinery is wearing away.

H-27. Solubility of Substances [7, 12, 15]

Have the class work in groups for this activity.

A series of six test tubes is set up, each containing 5 grams of a different solid such as calcium sulfate, aluminum sulfate, barium carbonate, sodium carbonate, and zinc oxide. After measuring the height of each substance, carefully add 5 ml of water to each tube, stopper the tubes, and shake them vigorously for 2 minutes. Let the tubes stand undisturbed for 15 minutes. Arrange the tubes in decreasing height of undissolved solid. Pupils then identify those which are quite soluble, slightly soluble, and insoluble (very low solubility).

Into each of two test tubes put 10 grams of potassium bromide. To one tube add 10 ml of cold water (less than 20°C) and to the other add 10 ml of hot water (more than 70°C). After 2 to 6 minutes, compare the amount of salt left undissolved.

H-28. Using a Solubility Table [5, 14, 15]

After showing pupils a solubility table, ask them to list the solubilities of substances such as sodium chloride, barium sulfate, potassium nitrate, copper sulfate, zinc hydroxide, and calcium sulfate.

Have the pupils decide which classes of compounds seem to be always soluble. (The ammonium, nitrate, sodium, and potassium compounds are always soluble.)

Some solubility curves and a table of solubilities in water are shown on pages 57 and 58.

H-29. Solubility Curve [7, 13, 15]

Each pupil places 1 gram of sodium chloride in a 250-ml beaker. From a graduated cylinder containing 100 ml of water, 1-ml portions are added to the salt at 30-second intervals, the mixture being constantly stirred until the salt completely dissolves. The temperature of the water is taken and the volume of water added is recorded. Data for the solubility curve plot for that temperature is calculated as follows:

$$100 \times \frac{\text{1 gram sodium chloride}}{\text{volume of water used}} = \text{amount needed to saturate 100 ml at the specified temperature.}$$

The procedure is repeated using water at higher or lower temperatures (at least 10° difference). Label temperature on the horizontal axis and grams per 100 ml of water on the vertical axis and plot the points. Since sodium chloride is just about as soluble in hot water as in cold water, a straight line graph most likely will be obtained. Using a more soluble salt introduces significant experimental error with the above method.

Solubility Curves

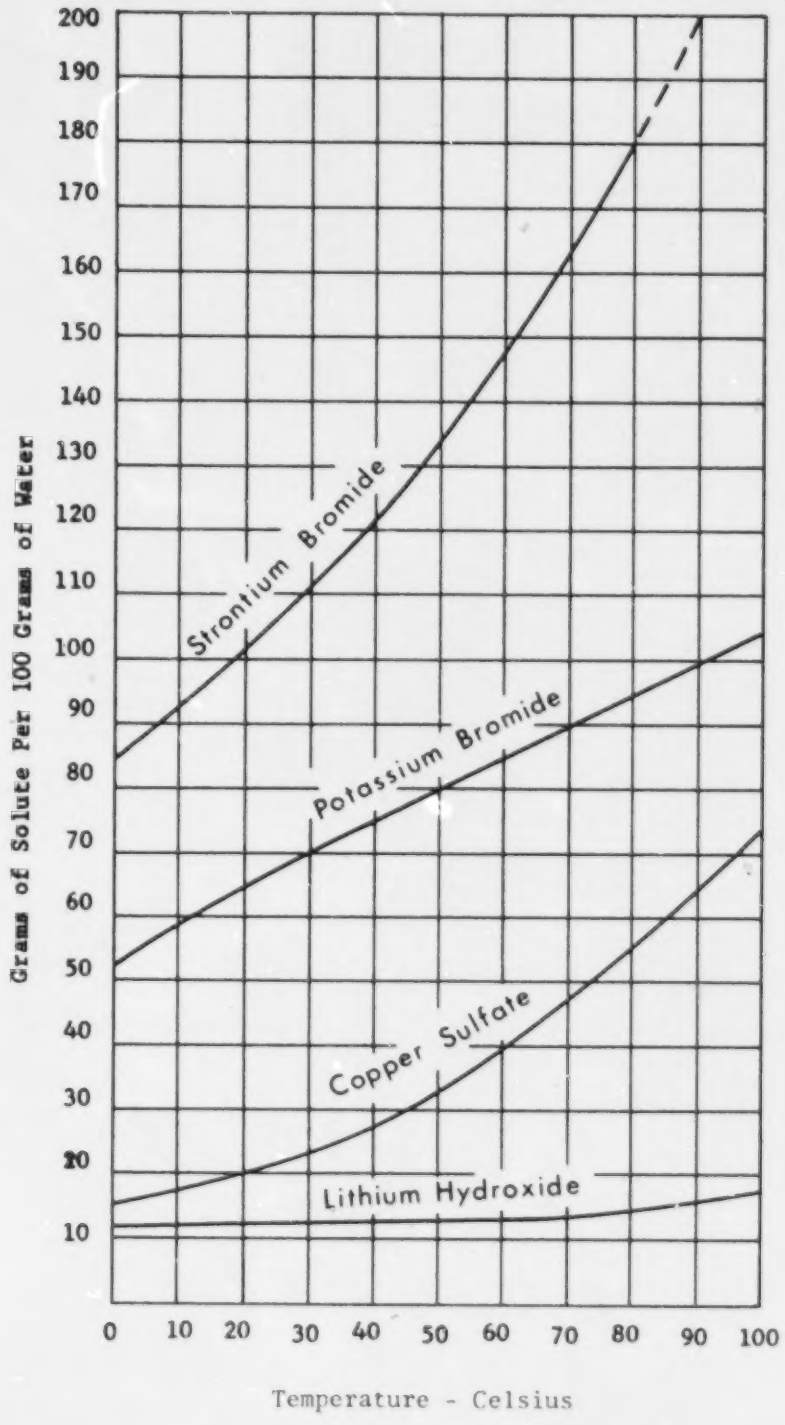


Table of Solubilities in Water

i-nearly insoluble ss-slightly soluble s-soluble d-decomposes n-not isolated	acetate	bromide	carbonate	chloride	hydroxide	iodide	nitrate	oxide	phosphate	sulfate	sulfide
aluminum	s	s	n	s	i	s	s	i	i	s	d
ammonium	s	s	s	s	s	s	s	n	s	s	s
barium	s	s	i	s	s	s	s	s	i	i	d
calcium	s	s	i	s	ss	s	s	ss	i	ss	d
copper II (cuprous)	s	s	i	s	i	d	s	i	i	s	i
iron II (ferrous)	s	s	i	s	i	s	s	i	i	s	i
iron III (ferric)	s	s	n	s	i	s	s	i	i	ss	d
lead	s	ss	i	ss	i	ss	s	i	i	i	i
magnesium	s	s	i	s	i	s	s	i	i	s	d
mercury I (mercurous)	ss	i	i	i	n	i	s	i	i	ss	i
mercury II (mercuric)	s	ss	i	s	i	i	s	i	i	d	i
potassium	s	s	s	s	s	s	s	s	s	s	s
silver	ss	i	i	i	n	i	s	i	i	ss	i
sodium	s	s	s	s	s	s	s	d	s	s	s
zinc	s	s	i	s	i	s	s	i	i	s	i

H-30. Solubility of Gases [2, 12, 14]

Place a bottle of soda water in ice water, chill, and open. Allow a second bottle to remain open at room temperature during the period. At the end of the period pour the contents of both into two glasses. Compare the amount of foam and the temperature of both solutions. The amount of foam is a rough approximation of the amount of gas that was dissolved in each.

Display two bottles of soda water that have been standing at room temperature for a period of time. Take the cap off one bottle. Ask the pupils to explain why gas escapes from the solution in the open bottle and why there are no bubbles of gas seen in the liquid in the capped bottle.

H-31. Separation by Filtering [4, 6]

Add water to each of the following combinations, stir, then filter. Retain the filtrate for use in Activity (H-32).

- iron filings and copper sulfate
- marble chips and magnesium sulfate
- sulfur pieces and potassium bromide

Mix equal volumes of 0.2 Molar sodium carbonate and 0.1 Molar barium chloride solutions. Filter to obtain the precipitate.

Support a large glass funnel on a ring stand. Fill the funnel by adding in order, clean pebbles, a 5-cm layer of coarse sand, and a 10-cm layer of very fine sand on top. Pour some muddy water through this filter unit. Remind pupils that while many solid impurities were removed, the water is not safe to drink.

H-32. Obtaining Solutes from Solution [4, 6]

Use a water bath to evaporate the filtrates obtained in Activity (H-31) in order to recover the soluble components of the mixtures.

Pour some saturated copper sulfate solution on a watch glass and set it on a window sill over the weekend.

Use the heat from a light bulb suspended over a watch glass containing a magnesium sulfate solution.

Evaporate equal volumes of water from various sources. If the residue chars on further heating, organic matter was present; if charring doesn't occur, dissolved minerals were present.

H-33. Changing Phase [1, 2]

Melt a few grams of Wood's metal in a Pyrex test tube. Pour the liquid metal into a lightly greased dish or on the demonstration desk. Compare the properties of the alloy as it changes from a liquid and back to a solid. Wood's metal can be obtained from most chemical supply houses.

H-34. Varying Degrees of Chemical Activity [1, 6, 14]

TEACHER DEMONSTRATION ONLY. Show that elements and compounds have varying degrees of chemical activity.

Using a copper strip, an iron nail, an aluminum strip, a magnesium ribbon, calcium metal turnings, charcoal lumps, and copper sulfate crystals, demonstrate the following reactions and compare the results. In no case use more than one small piece of calcium turnings.

(a) Ability to Combine with Oxygen. Heat a small piece of each substance in the oxidizing flame. Use long handle tongs to hold the substances as they are heated. Note the rate of reaction.

Caution pupils not to look directly at the flame in which the object is being heated. Vigorous reactions, such as those with magnesium or calcium can produce harmful rays which may injure the eyes.

(b) Ability to Combine with Water. Drop a piece of each substance into a few milliliters of cold water. Note any evidence of a chemical action.

(c) Combining with an Acid. Put a tiny piece of each material into a few milliliters of dilute hydrochloric acid. Note any evidence of chemical action.

(d) Reacting with Bases. Add a tiny piece of each material into a few milliliters of dilute sodium hydroxide solution.

(e) Comparing Results. Tabulate the results of each of the above reactions so that pupils may compare the differing degree, if any, with which each substance reacts with oxygen, water, an acid, and a base.

H-35. Recognizing Chemical Changes [10, 9]

Ask pupils to summarize the ways in which chemical changes were recognized after performing the following exercises.

- Drop chlorine bleach on colored construction paper, nylon hose, and cloth. (Pretest to be sure bleaching can be observed. Some dyes bleach by reduction rather than oxidation.)
- Use solutions of copper sulfate, aluminum sulfate, cobalt nitrate, iron (II) sulfate, and potassium chloride. In the wells of a spot plate or small test tubes, place two to five drops of a solution, then add drops of sodium hydroxide solution until a change can be observed.
- Drop hydrochloric acid on small chunks or small amounts of sodium carbonate, aluminum carbonate, potassium carbonate, and sodium bicarbonate.

TEACHER DEMONSTRATION ONLY. Before class place one or two drops of concentrated hydrochloric acid in a gas bottle and cover it. In another gas bottle place one drop of concentrated ammonia solution (ammonium hydroxide). Have the pupils observe the color of gases in the two bottles. Remove the covers and quickly hold the bottle containing the hydrogen chloride gas, evaporated from the acid, mouth downward about 3 to 5 cm above the mouth of the ammonia bottle. Ask the pupils to tell how they recognized that a chemical change had occurred.

H-36. Relative Amounts of Chemicals Used [9, 12, 13, 17]

Put 5 ml of 0.1 Molar lead nitrate solution into each of four test tubes. Add 2 ml of 0.2 Molar potassium chloride solution to the first tube, 4 ml to the second, and 6 ml to the third, and 8 ml to the last. Filter the contents of each tube, and add a few drops of 0.1 Molar potassium dichromate solution to each filtrate. Ask the class to propose a reason why the dichromate solution does not produce the same results in all tubes. (The yellow precipitate shows the presence of unreacted lead nitrate while the silver nitrate indicates unreacted potassium chloride.) Ask the class whether the same results would be obtained if a different chemical were used. Have pupils test their hypothesis by repeating the procedures, substituting a 0.1 Molar silver nitrate solution for the dichromate solution. (Remind pupils that silver nitrate can burn the skin.)

H-37. Energy and Chemical Reactions [6, 7, 12, 14]

After performing the following activities, have the class discuss how energy was involved in each case.

- Light a laboratory burner, a candle, a liquid butane lighter, and a safety match.
- In a dark room, place a key on a piece of blueprint paper, then expose the combination to bright sunlight for a one-half hour. "Develop" the paper by placing it in a pan into which water is overflowing.
- Take a flash picture, then inspect the flash bulb. Compare the appearance of a used and an unused bulb.
- In a darkroom, arrange some solid objects on a piece of photographic print paper. Expose the paper with the objects to intense light for a few seconds, then develop the paper with the solution recommended by the manufacturer.
- Make a simple electric cell. Soak an unglazed clay cup in a solution of potassium nitrate overnight. Remove the cup and add 0.1 Molar zinc nitrate so the cup is about two-thirds full. Place a zinc electrode in the cup. In a 250-ml beaker, pour about 75 ml of 0.1 Molar copper sulfate solution and place a copper electrode in the beaker. Proceed with the next section of the activity.

- Set the cup with its electrode into the beaker of copper sulfate. Connect the positive side of a voltmeter, 0 to 1.5 volt range, to the copper electrode. Connect the zinc electrode to the other terminal of the voltmeter.

- Perform the electrolysis of water. Note the effect of gas formation as the battery is turned on or off.

- In a styrofoam cup, pour 25 ml of 2 Molar hydrochloric acid and in another cup pour 25 ml of 2 Molar sodium hydroxide. Take the temperature of each solution. While stirring, pour the base into the acid and record the highest temperature produced in the mixture.

H-38. Changing the Rate of a Reaction [11, 14, 16, 17]

Have pupils perform the following activities and on the bases of their observations state the factors which seem to change the reaction rate.

- Pour 10 ml of 7% hydrogen peroxide into a loosely stoppered test tube and leave the test tube in the light until a change occurs. Remove the stopper and sprinkle a few grains of manganese (IV) oxide (manganese dioxide) into the hydrogen peroxide. Compare the rates at which gas bubbles are produced. When the action stops, filter the contents of the tube. Ask whether the manganese (IV) oxide on the filter has been changed.

- Into each of two test tubes place 10 ml of 2 Molar hydrochloric acid. Into one tube put a piece of mossy zinc (tech. grade) and into the other an equal size piece of pure mossy zinc. Ask the class to propose a reason for the difference in observations. Have them develop an experiment to test their hypothesis and then form and evaluate this hypothesis. (Tech. grade Zn contains considerable carbon which affects the rate by electrochemical action.)

- Compare the rate of gas formation when pieces of zinc of equal size are dropped into equal volumes of cold and warm dilute acetic acid. The differences in acid temperature are produced by keeping one test tube containing the acid in an ice bath and the other test tube in a hot water bath.

- Drop a chunk of limestone (marble chip) into 1 Molar hydrochloric acid and another chip of the same size into an equal volume of dilute sulfuric acid. Have the class propose a reason for the difference in reaction rates. (Calcium sulfate is only slightly soluble and soon coats the chip preventing the calcium carbonate from coming in contact with the acid.)

- Pour 50 ml of white vinegar into each of two 250-ml beakers. Simultaneously add 2 grams of sodium carbonate crystals to one beaker and 2 grams of pulverized sodium carbonate to the other. Have the class propose a reason for the difference in reaction rates.

- Pour water in increasing amounts by 3 ml into 5 test tubes. While stirring, add 5 ml of 1 Molar hydrochloric acid to each test tube. In a sixth tube, pour 5 ml of the acid also. Put a small piece of equal size of zinc into each solution. Compare the rate at which gas is released.

H-39. Using a Catalyst in a Chemical Change [1, 17]

Put 10 ml of 3 percent hydrogen peroxide into a test tube. After a few minutes observe the bubbles which have appeared. Insert a glowing splint into the mouth of the test tube and note that little or no oxygen is given off.

Sprinkle about one-quarter teaspoonful of manganese dioxide into the hydrogen peroxide. Observe that the gas bubbles form faster. Test the gas with a glowing splint.

The use of any hydrogen peroxide stronger than 3 percent is not recommended. The 3 percent hydrogen peroxide is the solution used as an antiseptic and may be obtained at drugstores.

H-40. Effect of Concentration upon Velocity of Reaction [17, 14, 1]

Make a solution containing approximately 10 gm of sodium thiosulfate (hypo) in 100 ml of solution.

Into three 150-ml beakers, place 10.0 ml, 5.0 ml, and 2.5 ml of solution. In the second and third beakers add 5.0 ml and 7.5 ml of water respectively so that each beaker contains 10 ml of solution. Place the beakers over a sheet of lined paper on a table.

Add 10 ml of hydrochloric acid (approximately 1N) to each beaker. View the lines on the paper by looking down through the solution. White colloidal sulfur forms in each case. With a stop watch determine the times required for the mixture to become opaque enough to make the lines invisible. Pretest the experiment and adjust the normality of the HCl if the time of reaction is not satisfactory.

Traces of poisonous sulfur dioxide may be detected from this reaction. Be sure to have the room well-ventilated while the experiment is being done.

H-41. Effect of Temperature on Velocity of Reaction [12, 13, 14]

Use the procedure listed in (H-40) but keep the volume of hypo constant at 10 ml. Vary the temperature of the hypo. Graph the results.

To get the varying temperature of hypo, bring some of the solution to a boil. As the solution cools, the desired quantities of hypo solution can be removed at different temperatures.

H-42. Existence of Atoms and Molecules [1, 2, 9, 6]

A variety of demonstrations can provide evidence of the kinetic-molecular theory which in turn will increase the understanding of atoms. Some suggestions are given below.

(a) Saturate a small piece of cotton with ether or ammonia, and place it in a dish. After a short time, ask pupils to explain what occurs.

(b) Using a thistle tube, half-fill several colorless glass medicine bottles with a 4 percent solution of gelatin. Take care to keep the gelatin from touching the upper part of the bottles. When the gels have set, fill the remaining space of the several bottles with solutions of potassium chromate, cobalt nitrate or of soluble dyes such as magenta. Stopper the bottles. Stand some bottles upright, clamp others upside down, and lay some on their sides. Have pupils observe in what directions the color advances.

(c) Fill two similar beakers nearly full of water. The water in one should be at room temperature and in the other nearly at the boiling point. Support two thistle tubes upright so that the tubes extend down the center of each beaker to about one-eighth inch from the bottom. Drop a few crystals of potassium permanganate down each tube. Observe the differences in the rate of diffusion.

(d) A Brownian movement demonstration may prove effective.

H-43. Determining Relative Masses [7, 12, 13, 15]

Have the pupils zero adjust a double pan balance. (The poise is not used in this experiment.) On the left hand balance place a penny. Balance the mass of the penny by adding paper clips to the other side. Record the observations. Repeat the procedure to find the numbers of clips representing the masses of other coins. Select one coin (other than the heaviest or lightest) and let that be the standard which can be assigned any whole number. The relative masses of the other coins may be calculated:

$$\text{relative mass of coin} = \text{standard no.} \times \frac{\text{no. clips for the coin}}{\text{no. clips for the standard}}$$

Relate the activity to the relative atomic masses based upon Carbon taken at 12 units. Remind pupils that the relative masses were derived from combining masses not paper clips!

H-44. Likenesses and Differences of Isotopes [18]

Have pupils use an analogy to consider differences and similarities of isotopes. Imbed different numbers of lead pellets inside styrofoam balls so that the number of pellets cannot be determined. Have the pupils list the ways in which the balls are alike, then determine the masses of the balls. Ask the pupils to propose a reason for the differences in masses. Ask the pupils to account for the differences between the isotopes chlorine-35 and chlorine-37.

H-45. Constructing Atomic Models [19, 22, 18]

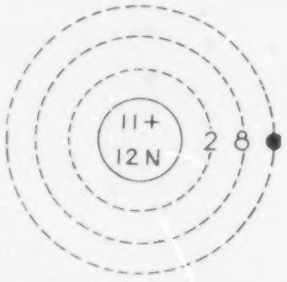
Using the chart below, pupils should be instructed how to prepare atomic model diagrams. Reference should be made at this point to the use of the Periodic Table in supplying the information given in the chart (see page 66).

Atomic Number	*Element*	*Symbol*	*Atomic Weight*	*Protons*	*Neutrons*	*Electron Arrangement*
1	Hydrogen	H	1	1	0	1
2	Helium	He	4	2	2	2
3	Lithium	Li	7	3	4	2,1
4	Beryllium	Be	9	4	5	2,2
5	Boron	B	11	5	6	2,3
6	Carbon	C	12	6	6	2,4
7	Nitrogen	N	14	7	7	2,5
8	Oxygen	O	16	8	8	2,6
9	Fluorine	F	19	9	10	2,7
10	Neon	Ne	20	10	10	2,8
11	Sodium	Na	23	11	12	2,8,1
12	Magnesium	Mg	24	12	12	2,8,2
13	Aluminum	Al	27	13	14	2,8,3
14	Silicon	Si	28	14	14	2,8,4
15	Phosphorus	P	31	15	16	2,8,5
16	Sulfur	S	32	16	16	2,8,6
17	Chlorine	Cl	35	17	18	2,8,7
18	Argon	Ar	40	18	22	2,8,8
19	Potassium	K	39	19	20	2,8,8,1
20	Calcium	Ca	40	20	20	2,8,8,2

Although the pupils should realize that the probable arrangement of the electrons is three-dimensional and rather vague, the two-dimensional diagram is generally satisfactory for elementary explanations and is easier to draw.

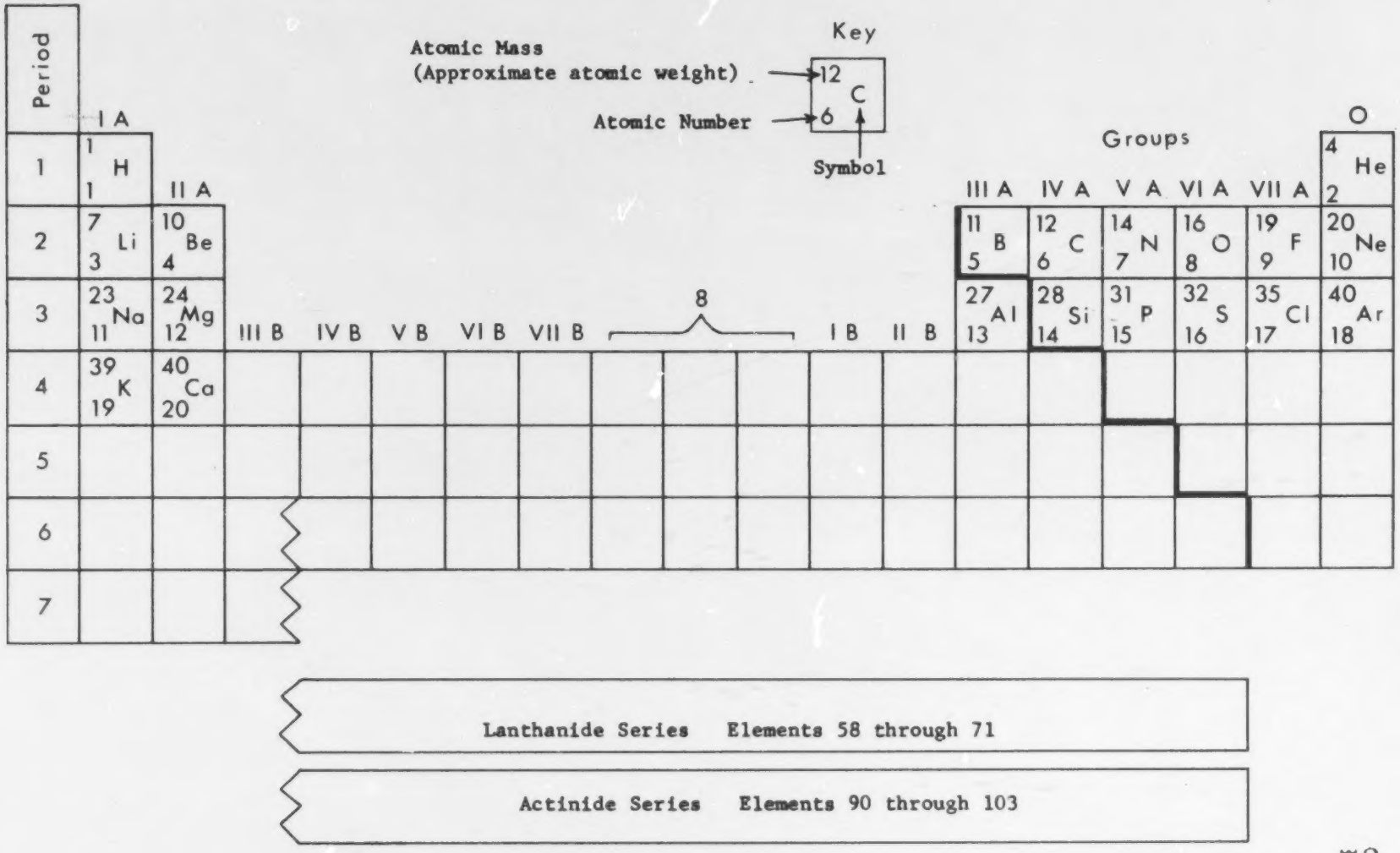
In order to prevent misinterpretation due to vague or crowded diagrams, a relatively uniform system of representation should be used. It is suggested that acceptable diagrams include the following features.

- A title beneath each diagram such as "Sodium Atom"
- The nucleus represented as a circle at least one-half inch in diameter
- In the nucleus the numbers of protons and neutrons clearly indicated by number and code such as protons (11+ or 11P) and neutrons (12N)



Sodium Atom

Periodic Table of the Elements



- The completed inner shells represented by dashed arcs or circles about one-fourth inch apart with numbers placed on each arc or circle to represent the number of electrons in each completed shell

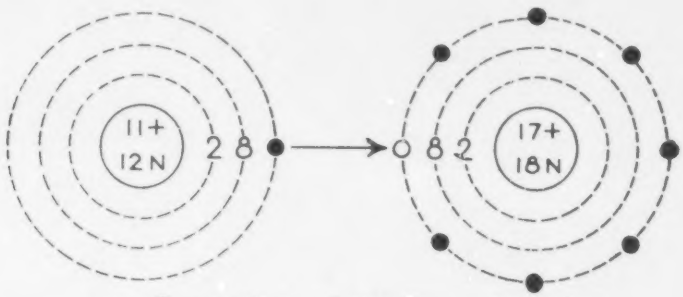
- The outer (valence) shell represented by a dashed circle with a radius about one-fourth inch larger than the radius of the preceding arc and electrons in the valence shell clearly shown as solid circles or by "e"

(a) Draw circles on a large sheet of paper. Paste or fasten the paper on tackboard or heavy cardboard. Use thumbtacks of three colors to represent the particles.

(b) For use with an overhead projector draw circles on a sheet of clear acetate with acetate-type ink. With a paper punch, punch out the particles from pieces of colored acetate or cellophane. Static electricity helps hold the particles in position. The entire set may be conveniently stored in an envelope.

H-46. Using Atomic Models [19, 16]

In diagraming the probable structure of ionic compounds, use an arrow to show the shifting of a valence electron to its new position. The shifting electron(s) may be shown as solid circles in the original position and as open circles in the new position. The diagram below represents sodium chloride.



Formation of Sodium Chloride

The electron arrangement of some of the other elements is shown in the chart in activity (H-45) under electron arrangement. Have pupils show how sulfur reacts with each of the following: sodium, magnesium, and aluminum.

H-47. The Periodic Table [13, 14]

Trace the outline and boxes of the Periodic Table shown on page 66 . Transfer the "skeleton" to a stencil or ditto master, and duplicate enough copies so that each pupil may have one. Refer to activity (H-45) for the electron arrangement of the first 20 elements.

Have the pupils follow the format shown in the key and fill in the atomic number, atomic mass, and electron configuration for each of the first 20 elements. Colored pencils may be used to shade in the areas where the metals and nonmetals are listed.

Show the class a completed Periodic Table. On the bases of their position in the table, have the pupils classify each of the following as a metal or a nonmetal: bromine, mercury, iron, barium, zinc, iodine, lead, and selenium.

H-48. Similarities Within a Group [14, 17, 21]

Arrange a display of the chlorides, hydroxides, nitrates, and carbonates of magnesium and calcium. Have the pupils observe whether the metals have formed similar compounds. Observations may include physical properties such as color, odor, and solubility. React the solutions of the compounds with dilute solutions of hydrochloric acid, sulfuric acid, sodium hydroxide, sodium bicarbonate, and sodium hydrogen phosphate to determine the chemical properties.

List the properties of sodium, and ask the class to propose an element having similar properties, then have the pupils do library research to test their hypothesis.

H-49. Some Direct Union Reactions [13, 14, 17]

Demonstrate some chemical reactions, then have the class decide, on the basis of their observations, how the reactions are similar.

- React fresh calcium oxide lumps with just enough water to cover the bottom of a beaker. Caution: spattering may occur and considerable heat will be liberated. Ask pupils to explain the appearance of steam.

- On a thick wire gauze under a hood, place a mixture containing 0.6 grams of zinc and 0.4 grams of powdered sulfur so the mixture forms a cone. Hold a lighted burner at arm's length and heat the top of the cone. A vigorous fireworks-like reaction will occur.

- In a test tube place a mixture of 5.6 grams of fine iron filings and 3.4 grams of sulfur. Gently heat the contents until the mixture becomes red-hot. Then remove it from the flame. Be careful to keep the burner away from the gaseous sulfur which may escape from the mouth of the test tube.

H-50. Recognizing Single Replacement Reactions [11, 12, 13]

After performing the activities below, have the pupils develop a statement which explains how a single replacement reaction can occur.

Place a piece of metal into saturated copper sulfate solution. Use a different metal such as lead, zinc, iron (nail), steel paper clip, and magnesium. Compare the time it takes for changes to occur in each test tube.

Ask the pupils what will happen if they repeat the above procedures substituting dilute lead nitrate for the copper sulfate. After the class has performed the activity using lead nitrate, have the pupils revise their hypothesis if necessary.

Into a test tube half-full of silver nitrate solution suspend a clean bare copper wire. Leave the setup stand undisturbed overnight. After the observations have been made the next day, collect the solid residue by filtering, dry it, then test its malleability. Evaporate the filtrate. Have the pupils state a hypothesis about the composition of the solid obtained by evaporation and check the validity of their hypothesis by an experiment which they develop.

H-51. Relative Activity of Metals [10, 12, 13, 16]

Ask the pupils to observe the reactions occurring when different metals are placed in a solution. The metals used may be copper, iron, lead, magnesium, and zinc. Solutions may include dilute hydrochloric or sulfuric acid, copper sulfate, iron (II) sulfate, lead (II) nitrate, magnesium sulfate, silver nitrate, and zinc nitrate. Other salt solutions may be used also. Point out that observing what doesn't occur often gives as much information as what does occur. Ask the pupils to give a possible explanation for what they observed. Ask the class how they would prove that iron was more active than silver. Ask them to test their hypothesis.

Have the pupils number the periods and groups and mark the zig-zag line running from boron to the bottom of period 6. With a crayon or colored pencil shade in the area between the period numbers and the zig-zag line. Label the area metals. The area between the zig-zag line and group 0 should be shaded with another color and be labeled nonmetals. A third color may be used to tint group 0 which is labeled inert gases.

H-52. Single Replacement Reactions [11, 1, 14, 6]

Give each pupil a piece of copper, zinc, lead, and iron nails. Also have available the following solutions: silver nitrate (2 grams per liter), copper (II) sulfate, iron chloride, and sodium chloride.

Instruct the pupils to place a clean piece of one of the metals into 5 to 10 ml of any one of the solutions in a test tube. After 5 minutes observe the metal and note any formation of a deposit or change in the color of the solution. Identify the products by writing a word equation. Check the activity chart shown below to determine if the reaction should occur.

Activity Series

Metals	Nonmetals
(K) Potassium	(F) Fluorine
(Ca) Calcium	(Cl) Chlorine
(Na) Sodium	(Br) Bromine
(Mg) Magnesium	(I) Iodine
(Al) Aluminum	

(Zn) Zinc
(Fe) Iron
(Ni) Nickel
(Sn) Tin
(Pb) Lead
(Cu) Copper
(H) Hydrogen
(Hg) Mercury
(Ag) Silver
(Au) Gold

For example, if iron and copper sulfate are placed together, there will be a reaction which forms iron sulfate and copper. Iron is more active than the copper and replaces the copper. If, however, copper and iron(II) chloride are placed together, no reaction will occur since copper is less active than iron.

Pupils may repeat the procedure using different combinations of metal and solution. UNDER NO CIRCUMSTANCES should pupils be given potassium, calcium, or sodium for any laboratory activity. Require the identification of products to discourage pupils from "playing" with chemicals.

H-53. Ionic Exchange Reactions [2, 14, 19]

Make solutions of various compounds such as copper (II) sulfate, barium chloride, potassium iodine, sodium bromide, aluminum sulfate, potassium dichromate, iron (II) sulfate, silver nitrate (2 grams per liter), and sodium hydroxide (4 grams per liter).

Indicate to the pupils that they can combine any of the solutions they wish as long as they do not use more than 5 ml of any solution and use no more than two solutions at a time. Have pupils decide whether or not a reaction occurred.

Require the identification of any precipitate. Pupils should write a word equation and then check the solubility of each product by referring to the table of solubilities shown on page 58 . More advanced pupils may refer to a handbook for solubility of any substance not listed on the table of solubilities.

H-54. Decomposition Reactions [9, 10, 11, 13]

(a) Have pupils place some copper (II) sulfate crystals in a test tube held so that the mouth of the tube is slanted slightly downward, and gently heat the crystals until the crystals have changed their appearance. Ask the pupils to account for the change and how they could reform the original crystals. Add a few drops of water and account for the apparent change.

(b) Demonstrate the electrolysis of water and identify the products.

(c) Have the pupils perform the electrolysis of copper (II) chloride solution in a U-tube equipped with carbon electrodes mounted in

one-hole stoppers which fit into the U-tube. Then test the solution in the arm to which the + electrode is connected with a piece of red or blue litmus and observe the result. Ask the class why the color of the litmus changed. Was the change due to the solution of copper (II) chloride or to the product formed at the electrode? Have the pupils check their hypothesis by an appropriate test.

Ask the class to study the method and results of the above reactions to find out how the reactions are all alike.

H-55. Frequently Used Chemical Symbols

Pupils should learn the following symbols which are most frequently used in the science laboratory.

aluminum	Al	nitrogen	N
calcium	Ca	oxygen	O
chlorine	Cl	phosphorous	P
copper	Cu	potassium	K
hydrogen	H	sulfur	S
iodine	I	tin	Sn
magnesium	Mg	uranium	U
mercury	Hg	zinc	Zn

Do not require pupils to learn the name of the symbols for all of the elements. Limit the requirement to the names and symbols listed above.

H-56. Information From a Formula [19]

(a) In containers labeled only with the formula of the contents, display a number of substances which have been used during the course. Ask the pupils to name the elements represented by each symbol in the formula and to state the number of atoms of each element as shown by the subscripts.

(b) Have pupils use an atomic model diagram to prove whether or not the formula given for a specific compound is correct. Some formulas given to the students might include: CaS, $CaCl_2$, LiF_2, Li_2O, Al_2O_3, and any other combination of binary, ionic compounds formed by combinations of the first 20 elements. Include a few incorrect formulas among those given.

(c) Provide the pupils with a list of formulas containing brackets or parentheses and ask them to indicate what elements are present and how many atoms of each element are represented.

H-57. Information From an Equation [20]

Write the equation $2Mg + O_2 \longrightarrow 2MgO$ on the chalkboard. Have the pupils list all of the information that can be obtained from this equation. The list should include the names of reactants and products, what elements were present in each formula, how many atoms of each element were involved in the reaction, and how many molecules of oxygen were used.

H-58. Gram-Atomic Masses in a Sample [7, 5, 12]

Provide the pupils with a sample of an element such as sulfur or iron filings. Ask the pupils to predict how many gram-atomic masses are present in the given sample, then the pupils should determine the mass of the sample and calculate the number of gram-atomic masses present:

$$\frac{\text{mass of the sample}}{\text{grams per gram-atomic mass}} = \text{no. of gram-atomic masses present}$$

Repeat the procedure using metals such as aluminum foil and lead pellets. Have the students measure out 0.1 and 0.5 gram-atomic masses of powdered sulfur or iron filings. If pupils seem confused because the volume of the samples are not consistent, remind them that they are using mass relationships which depend upon the relative masses of the atoms, not on the volume of the samples.

H-59. Determining the Number of Gram-Formula Masses Present [5, 12]

Repeat the procedures in (H-58) to determine the masses of given samples of some salts. After calculating the gram-formula masses, the pupils calculate the number of gram-formula masses present in the sample by using the following formula:

$$\frac{\text{mass of the sample}}{\text{grams per gram-formula mass}} = \text{no. of gram-formula masses present}$$

Ask pupils to measure out 0.1 gram-formula mass of magnesium sulfate or copper (II) sulfate. By dissolving the 0.1 gram-formula mass samples in 100 ml of water one can prepare 1 Molar solutions which can be used in other activities.

H-60. Properties of Acids and Bases [9, 14, 17]

- Note the color change produced on various indicators by acids and bases. Some indicators which may be used are litmus, methyl, orange, brom phenol blue, phenolphthalein, and universal indicator.

  React dilute acids with metals. Review the activity series of the metals for single replacement reactions.

- Observe the effect acids and bases have on the color of juices obtained from fruits and berries. State a hypothesis accounting for this change.

- Determine which rock samples contain carbonates by dropping hydrochloric acid on each.

- Use litmus to see which household cleaners in water produce the characteristics of a basic solution.

- Place a small amount of ground lean meat in a test tube and add hydrochloric acid. Let the tube and contents sit overnight and observe. Relate this activity to digestion.

- Make soap by reacting a fat and sodium hydroxide. (Relate the activity to cleaning clogged drains.)

- Place some barium hydroxide powder on a watch glass and let it sit in air overnight. Add a few drops of hydrochloric acid to the powder and ask pupils to explain why effervescence occurs.

H-61. Neutralization [19, 21, 22]

Have pupils place 5 to 10 ml of solution hydroxide (1 Molar) into a small beaker and add 2 or 3 drops of phenolphthalein indicator. While stirring, add 1 Molar hydrochloric acid dropwise until the indicator just turns colorless. Ask the class whether or not the same results would be obtained if the base were added to the acid. Have them test their hypothesis. The acid may be transferred by a long dropper from a graduated cylinder with the initial and final volumes of acid recorded or delivered from a burette. In the latter case, have pupils practice the technique with water in the burette.

Find the volume of 1 Molar hydrochloric acid needed to just neutralize 10 ml of 1 Molar sodium hydroxide. Give some students a 2 Molar acid and ask the class to account for the difference in volume needed to neutralize the base.

Have the students determine whether or not the type of acid used affects the neutralization of 10 ml of 1 Molar sodium hydroxide solution.

Neutralize phosphoric acid with sodium hydroxide solution using the procedures given above.

Pupils may compare the use of different indicators to determine the end point.

H-62. Preparation of Bases [11, 6, 14]

Review the action of calcium oxide and water to produce calcium hydroxide: $CaO + H_2O \longrightarrow Ca(OH)_2$

To solutions of various metallic salts, add sodium hydroxide solution drop by drop to produce a base by an exchange reaction. Have pupils use a solubility chart to determine whether any of the precipitates may be a hydroxide. (If an excess of sodium hydroxide is added to some salt solutions, the metallic hydroxide may react with the excess sodium hydroxide to form another salt. These reactions can be detected since there tends to be a change in the appearance or solubility of the hydroxide.)

H-63. Some Properties of Salts [13]

Provide pupils with 10-gram samples of solid copper (II) sulfate, barium chloride, cobalt nitrate, potassium bromide, and iron(II) sulfate. Pupils

compare the salt's color and solubility in 50 ml of water, and react 10 ml of each solution with each of the following solutions: sodium carbonate, calcium chloride, lead nitrate, and sodium hydroxide. Ask the pupils to make a list of the general properties of salts.

H-64. Comparison of the Strengths of Vinegars [6, 14, 3, 13]

This exercise will require two or three periods since pupils will need to develop new techniques. Burettes filled with water can be used during practice sessions.

The relative strengths of vinegars (amounts of acetic acid) can be determined by comparing the amount of vinegar required to neutralize 10 ml of a solution of sodium hydroxide.

Obtain several brands of vinegar. If necessary, filter the vinegar to remove any sediment.

Prepare a sufficient amount of sodium hydroxide solution so that there will be at least 150 ml of solution for each pupil using a burette for titration. The concentration should be 20 grams per liter. To 1 liter of distilled water add 20 grams of sodium hydroxide and stir. Since solid hydroxide absorbs moisture from the air, the weighing should be done rapidly when preparing the solution. However, if no calculations of the exact strength of vinegar are to be made, the amount of sodium hydroxide added need not be exact provided each pupil uses the same concentration for each of his titrations.

Pupils should be warned that sodium hydroxide is caustic and can "burn" the body (eye). Teachers may wish to distribute the measured amounts of solution to the pupils.

A pair of burettes are necessary for this procedure. These should be thoroughly cleansed and rinsed; first with distilled water and then with about 10 ml of the solution with which the burette is to be filled. Fill one burette with the vinegar to be tested and the other with the base. The tip of each burette should be filled by opening the stopcock momentarily and letting a small amount of the solution flow through the tip into the beaker containing that solution. Add sufficient solution to the burette, and adjust the liquid level so that the bottom of the meniscus is on the zero mark of the burette.

Remove 2 ml of base from the burette into a clean 150-ml beaker; add two drops of phenolphthalein. Place the beaker under the burette containing vinegar. Add vinegar dropwise with constant stirring until the color disappears. A white piece of paper under the beaker will make the color change more perceptible. When the addition of one drop of acid makes the color disappear, the end point of the titration has been reached. Read the burette to determine the amount of vinegar used.

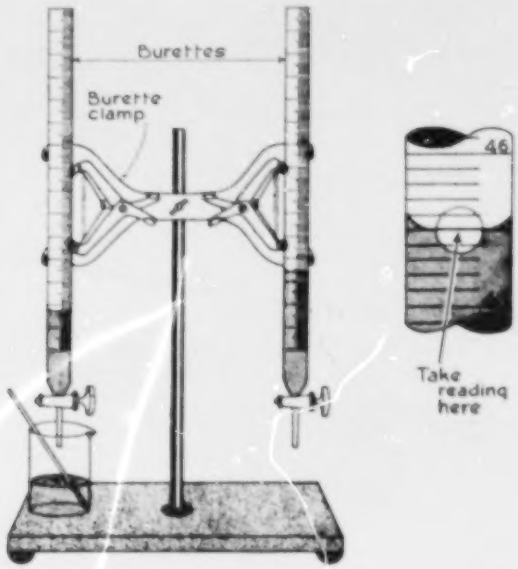
The teacher must judge how far to go in making quantitative comparisons of the acid concentrations.

Some pupils may select the strongest vinegar on the basis of the amount of vinegar required to neutralize the 20 ml of base. (The smaller the volume of vinegar used, the stronger the acid.)

Pupils of better ability can calculate the mass of acetic acid per 100 ml in each sample of vinegar. Twenty milliliters of sodium hydroxide contain 0.4 grams of sodium hydroxide. The combining mass ratio is 2 parts of sodium hydroxide to 3 parts of acetic acid. Therefore, 0.4 grams of sodium hydroxide will be neutralized by 0.6 grams of acetic acid. The 0.6 grams of acid are contained in the volume of vinegar used in the neutralization. To find the mass of acid in 100 ml of solution, divide 0.6 grams by its volume and multiply by 100.

A comparison of the mass of acetic acid per 100 ml for each acid shows which is the strongest one. (The greater the mass per 100 ml, the stronger the acid.)

For best results the titration should be repeated and an average taken for three readings. Caution: Sodium hydroxide should not be left in burettes for any length of time as the base can react with glassware.



H-65. Making Suspensions and Studying Their Properties [1, 6, 12]

Mix dirt with varying amounts of water and note how long it takes each to settle. Have pupils vary other factors such as temperature, presence of pollutants, etc. Data may be graphed.

See if light can be seen passing through some "chocolate milk" which is made by mixing cocoa and milk. Let the mixture stand to see if it settles. Vary the activity by having some pupils use more cocoa. Record results.

Add clay to water to make a suspension. Vary the amounts of clay in the same volumes of water to see if the settling rate varies with the amount of suspended material. Shake each sample to resuspend the clay, and retime the settling.

H-66. Properties of Solutions [1, 2]

Have pupils observe the settling rate, ability to be filtered, and ability to transmit light of several colored solutions such as dilute cobalt nitrate, iron (III) chloride, nickel sulfate, potassium sulfate, potassium dichromate, and potassium permanganate. The solutions should be diluted so that the colors of any solution are visible but faint. Instruct the class to make a solution of copper (II) sulfate. (Do not give any directions other than to record the mass of copper sulfate and the volume of water used.) Have the class compare the observations made and the quantities used to make the solutions. Ask the class to develop a list of properties of solution and to define a solution.

H-67. Concentrations of Solutions [6, 1, 9]

Display a number of solutions for which the concentrations are listed on the labels. The concentrations should include dilute, concentrated, number of grams per liter, number of grams per milliliter saturated, and unsaturated. A molar solution should be labeled with 1M and the gram-formula mass per liter. Ask the pupils to explain how one can tell how much solute has been dissolved in the solvent.

Place five crystals of hypo (sodium thiosulfate) in 1 ml of water. Shake until the crystals have been dissolved, then repeat the process as many times as necessary until no more crystals can be dissolved. Add 10 grams of crystals to the solution just made and heat it until it boils. Cool the solution under running cold water but be careful not to shake or agitate the solution while it is cooling. Have the class discuss the relative amounts used and compare the dissolving conditions. Ask the pupils to add one crystal of hypo to the supersaturated solution and account for what they observe. Discuss the reason for the apparent instability of one of the solutions.

H-68. Factors Affecting Solubility [1, 10]

- Obtain a large crystal and an equal weight of tiny crystals of potassium dichromate or copper sulfate. Place two petri dishes containing equal volumes of water on the overhead projector. Into one dish drop the large crystal, into the other dish the small crystals. Observe the action in each dish for several minutes. Compare the color of the solution in the dishes. Ask the pupils for a possible explanation of the results. Usually they will suggest that more water comes in contact with the small crystals than with the large crystal.

- To 100 grams of potassium bromide in a beaker add 100 ml of distilled or deionized water. Stir until no further solution seems to take place. Observe the temperature. Ask the pupils to

determine the weight of the salt now in solution by referring to solubility tables. Have them predict how warm the solution would have to be to dissolve all of the salt. While stirring frequently, heat the solution to the selected temperature. Place a little of the warm solution in a test tube or in a petri dish to cool. With an overhead projector show what occurs within the petri dish as the solution cools. Ask the pupils to explain why crystals appear in the dish as the temperature of the solution decreases.

- Place a bottle of soda water in ice water, chill, and open. Allow a second bottle to remain open at room temperature during the period. At the end of the period pour the contents of both into two glasses. Compare the amount of foam and the temperature of both solutions. The amount of foam is a rough approximation of the amount of gas that was dissolved in each solution.

- Display two bottles of soda water that have been standing at room temperature for a period of time. Take the cap off one bottle. Ask the pupils to explain why gas escapes from the solution in the open bottle and why there are no bubbles of gas seen in the liquid in the capped bottle.

H-69. Introducing Solubility Curves [5, 14, 1]

Before class time prepare a solubility curve to be shown to the pupils. The curve may be drawn on the blackboard, mimeographed, or printed on a transparency for projection on an overhead projector.

Introduce the pupils to the graph by telling them that a solubility curve is a mathematical picture or story about how the solubility of a substance changes as the temperature changes. Since the "story" concerns only two variable characters, the amount of solute and the temperature, all the other factors such as the amount of water or other solvent used must remain the same.

Have the pupils study the curve for a few minutes. Ask the pupils to tell what information they have found about the axes. Review with them the way in which to read a graph. Read crosswise between the mass of solute axis and the curve; read up and down between the curve and the temperature axis.

Have the pupils determine how much solute is needed to saturate 100 grams of solvent at different temperatures.

Remind the pupils that if a solution does not contain as many grams of solute as the curve indicates it can, the solution is unsaturated. More solute must be added to saturate it. On the other hand, a supersaturated solution would have more grams of solute dissolved than the curve indicates it can at the existing temperature. The difference between the amount of solute present in a supersaturated solution and the amount the curve shows is the weight of crystals that would come out of the supersaturated solution as the solution returned to the saturated condition.

H-70. Plotting a Solubility Curve [13, 14, 15]

Pupils can plot a solubility curve from data obtained from references or from experimental data they obtain.

(a) Using Data from Resources. Give the pupils the data shown at the right which is needed to plot the solubility curve of copper sulfate. The curves that illustrate these data can be found on page 57. Have the pupils set up the temperature scale on the horizontal scale. The data are plotted and a curve drawn connecting the points. Unless a straight line curve is clearly indicated, do not have the pupils use a ruler to connect consecutive points. A line should be lightly drawn, free hand, connecting the points, smoothed out, and then darkened to give a satisfactory curve. Label the curve with the name of the solute.

It may be wise to review the plotting of points. Stress the need to check each point plotted. Read down the line from a point to check the temperature; read across the line to check the mass of solute.

Have pupils plot the other curves for which data are given at the right. The curves can be used to solve problems found in the activities of appendix (H-71).

For enrichment, have the pupils consult one of the handbooks for data for the solubility of other substances.

Solubility of certain anhydrous compounds in grams per 100 grams of H_2O

Temperature - degrees Celsius	LiOH	$CuSO_4$	KBr	$SrBr_2$
100°	17.5	75.4	104.0	222.5
90°	- - -	- - -	99.2	- - -
80°	15.3	55	95.0	181.8
70°	- - -	- - -	90.0	- - -
60°	13.8	40	85.5	150
50°	13.3	33.3	80.2	135.8
40°	13	28.5	75.5	123.2
30°	12.9	25	70.6	111.9
20°	12.8	20.7	65.2	102.4
10°	12.7	17.4	59.5	93
0°	12.7	14.3	53.5	85.2

(b) Using Experimental Data. This activity should be restricted to those pupils showing considerable laboratory skills and an awareness of experimental errors. All but the better students may find the exercise frustrating and meaningless.

The data must be obtained by a team of students, each one being responsible for a specific temperature. The exercise requires too much time for one pupil to obtain all of the necessary data.

While each pupil is massing an empty, clean, dry evaporating dish, the teacher should prepare a hot, saturated solution of potassium chloride.

Before pouring the saturated solution into an evaporating dish, read the temperature of the solution; and then pour in enough solution to half-fill the dish. As the solution cools, give other pupils samples of the solution. Make sure each pupil knows the temperature at which he receives the solution.

The pupil weighs (masses) the dish and its contents carefully. Then, place the dish on fireproof material. Arrange a heat lamp so that the heat is directed down on the solution. Heat the solution to evaporate it. (If a bunsen burner is used, the solution is apt to spatter and lose some of the contents to the environment.) Weigh the dish and the dry salt. Perform the following calculations:

1. Mass of solution used =
   [mass of dish + solution] - mass of empty dish

2. Mass of water in the solution =
   [mass of dish = solution] - [mass of dish + dry salt]

3. Mass of salt = [mass of dish + salt] - mass of empty dish]

Since the solubility curve is set up on a basis of mass per 100 grams of water, convert the data obtained to the 100 grams of water basis. A simple calculation can be used:

$$100 \text{ grams} \times \frac{\text{mass of salt}}{\text{mass of } H_2O} = \text{mass of KCl (plotted)}$$

Have the pupils record on a chart the values calculated and the temperature at which they were obtained. The class can plot the data points and then compare the location of the data with the curve obtained by using the values obtained from a handbook. Do not try to connect the points obtained by the pupils. Generally about half will be too high and the others too low in value to match the standard curve.

The sources of error are great. The limitations of equipment, the inexperience of the pupils, the presence of some potassium chloride crystals in the saturated solution, spattering or spilling contribute to the obtaining of points which do not fall in the "real" solubility graph. An understanding of experimental error is the major outcome of this exercise.

H-71. Using the Solubility Curves to Solve Problems [14]

Pupils can use the solubility curves plotted in appendix (H-70) to determine a number of factors. Some sample problems follow:

(1) *Can 10 grams of anhydrous copper* (II) *sulfate saturate 100 grams of water at 40°C ? If not, how much more solute is necessary?*

From the $CuSO_4$ curve:

28.5 g are required to saturate 100 g of water at 40°C.
10.0 g given (would make unsaturated solution)
18.5 g additional $CuSO_4$ needed to make a saturated solution.

(2) *At what temperature will 25 grams of anhydrous copper* (II) *sulfate saturate 100 grams of water?*

Look across 25-gram solute line to $CuSO_4$ curve, then down to the temperature axis. The temperature is 30°C.

25 g of $CuSO_4$ will saturate 100 g of H_2O at 30°C.

(3) *At what temperature will 25 grams of copper* (II) *sulfate saturate 80 grams of water?*

This is a more difficult problem. In order to use the curve, it is necessary to determine how much copper sulfate would be in 100 grams of water if both solutions are of the same concentration.

Use the equation:

$$\frac{\text{25 grams of salt}}{\text{80 grams of water}} = \frac{X \text{ grams of salt}}{\text{100 grams of water}}$$

Therefore, 31.25 grams of the salt are required. The temperature shown on the graph is approximately 45°C.

(4) *A solution of strontium bromide is saturated at 20°C. How much additional solute must be added to keep the solution saturated as the temperature increases to 30°C?*

Using $SrBr_2$ curves:

at 30° 112 g needed
at 20° 102 g needed
10 g must be added to keep the solution saturated. Pupils will read approximate values in most cases.

(5) *At 100°C , 100 grams of water is saturated with lithium hydroxide. The solution is cooled to 0°. How many grams of solid lithium hydroxide will settle out of the solution? What kind of solution is left?*

At 100° 17.5 g LiOH required for saturation
At 0° 12.7 g LiOH required for saturation
4.8 g will crystalize out of solution

The remaining solution is saturated.

The teacher can make variations on the above problems by using different solutes, different amounts of water, different temperatures and by starting with supersaturated or dilute solutions which should be changed to saturated ones at a specific temperature. Better pupils should be able to work with values which require estimations.

Allow one or two points leeway from the exact graph value in determining correct answers to any of the above values. The curves that the pupils plotted are not exact.

H-72. Preparation of Colloids [6, 1, 17]

Have the pupils prepare some colloidal dispersions such as those suggested below.

- Add a few drops of milk or milk of magnesia to distilled water and stir.
- To a solution of hypo (1 gram per 100 ml), add 5 to 10 drops of 6 Molar hydrochloric acid and stir. After a few minutes, colloidal sulfur should be formed. Filter to obtain the colloidal dispersion.
- Place 1 gram of gelatin in 50 ml of water. After a few minutes add 100 ml of boiling water. Let the mixture cool.
- Stir some wave-setting gel or lotion into some water.

After the colloids have been prepared, have the pupils observe the appearance of the collodial dispersion, test the way in which light travels through the dispersion, and determine whether or not filtering separates out the dispersed material. The Tyndall effect can be easily seen if the beaker containing the colloid is placed inside a heavy paper bag which has a small hole position near the side of the beaker. Use a pencil type flashlight as the source of light applied at the hole in the bag.

Have pupils list the properties of colloids and compare them with those observed for solutions and suspensions.

Use the Tyndall effect to test a number of common substances to see whether they are colloids. If necessary, dilute or suspend the material in water. Some substances which may be used are ketchup, a creamy salad dressing, a lotion, jello or gelatine dessert, and a soap or detergent.

H-73. Identification of Organic Substances [2, 1]

Have the pupils list the properties of some organic substances by observing the following activities.

- Evaporate 50 ml of water samples from a pond, stream, river, and local water supply. Continue to heat the residue for 2 or 3 minutes.
- Place about five crystals of sugar in a test tube and heat gently for a few minutes. (Use the hood if larger quantities are heated.) Repeat the procedure using sodium chloride crystals.
- Under the hood, the teacher may wish to heat small pieces of fabric or yarn in test tubes.
- Heat a 1-cubic centimeter piece of bacon in an evaporating dish until it chars. Repeat the process using pieces of bread, crackers, or other food. A few leaves of celery may also be used.
- Demonstrate the electrical conductivity of acid, base, and salt solutions. Then test the conductivity of some organic compounds such as sugar solution, alcohol, and glycerine. Other properties such as odor, phase, melting point, and solubility may be discussed by the teacher.

H-74. Differences Between Organic and Inorganic Compounds [1]

Many differences can be illustrated by referring to the *Handbook of Chemistry and Physics*, to Lange's *Handbook of Chemistry*, or to similar lists of physical constants.

(a) Number of compounds. Have pupils determine the number of pages devoted to listing inorganic and organic compounds.

(b) Decomposition. Compare a few pages at random to note whether more organic or inorganic compounds decompose.

(c) Solubility. Compare a few pages at random to note which type is more soluble in water.

H-75. Ionic and Covalent Bonds' Analogy [17]

The differences between ionic and covalent bonds may be shown by an analogy. Assume that one person gives a dollar bill to another. Although one has lost a dollar and the other has gained a dollar, each may proceed and the second person may use the dollar, in any way he or she wishes. This "represents" the *ionic* bond. Now assume that one person hands one end of a dollar bill to another but holds the other end. Although each person thinks he or she has the dollar, it is really shared and nothing can be done with it unless both agree. This "represents" the *covalent* bond.

H-76. Esterification [1, 2]

Esters are most frequently prepared by the reaction of alcohols and acids in the presence of concentrated sulfuric acid. When the acid used is an organic acid, many of the esters resulting from such reactions have odors characteristic of certain fragrances produced by plants. It is possible to prepare various esters in the laboratory which pupils may readily recognize.

APPENDIX - LEARNING ACTIVITIES

In separate test tubes place the amounts of acids and alcohols as indicated below. Add a few drops of concentrated sulfuric acid to each tube. Heat by placing the tube in a beaker of heated water. Try to identify the fragrance.

- Five ml of methyl alcohol and 2 g of salicylic acid (wintergreen fragrance results).
- Five ml of ethyl alcohol and 5 ml of butyric acid (pineapple odor results).
- Five ml of n-amyl alcohol and 5 ml of acetic acid (banana fragrance results).